### Periodic Table of Elements

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<td>Li, Be, B, C, N, O, F, Ne</td>
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BASIC CHEMISTRY

Fifth Edition

Global Edition

KAREN TIMBERLAKE

WILLIAM TIMBERLAKE
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KAREN TIMBERLAKE is Professor Emerita of Chemistry at Los Angeles Valley College, where she taught chemistry for allied health and preparatory chemistry for 36 years. She received her bachelor’s degree in chemistry from the University of Washington and her master’s degree in biochemistry from the University of California at Los Angeles.

Professor Timberlake has been writing chemistry textbooks for 40 years. During that time, her name has become associated with the strategic use of pedagogical tools that promote student success in chemistry and the application of chemistry to real-life situations. More than one million students have learned chemistry using texts, laboratory manuals, and study guides written by Karen Timberlake. In addition to Basic Chemistry, fifth edition, she is also the author of General, Organic, and Biological Chemistry: Structures of Life, fifth edition, with the accompanying Study Guide, and Chemistry: An Introduction to General, Organic, and Biological Chemistry, twelfth edition, with the accompanying Study Guide, and Selected Solutions Manual, Laboratory Manual, and Essential Laboratory Manual.

Professor Timberlake belongs to numerous scientific and educational organizations including the American Chemical Society (ACS) and the National Science Teachers Association (NSTA). She has been the Western Regional Winner of Excellence in College Chemistry Teaching Award given by the Chemical Manufacturers Association. She received the McGuffey Award in Physical Sciences from the Textbook Authors Association for her textbook Chemistry: An Introduction to General, Organic, and Biological Chemistry, eighth edition. She received the “Texty” Textbook Excellence Award from the Textbook Authors Association for the first edition of Basic Chemistry. She has participated in education grants for science teaching including the Los Angeles Collaborative for Teaching Excellence (LACTE) and a Title III grant at her college. She speaks at conferences and educational meetings on the use of student-centered teaching methods in chemistry to promote the learning success of students.

Her husband, William Timberlake, who is the coauthor of this text, is Professor Emeritus of Chemistry at Los Angeles Harbor College, where he taught preparatory and organic chemistry for 36 years. He received his bachelor’s degree in chemistry from Carnegie Mellon University and his master’s degree in organic chemistry from the University of California at Los Angeles.

When the Professors Timberlake are not writing textbooks, they relax by playing tennis, ballroom dancing, hiking, traveling, trying new restaurants, cooking, and taking care of their grandchildren, Daniel and Emily.

DEDICATION

- Our son, John, daughter-in-law, Cindy, grandson, Daniel, and granddaughter, Emily, for the precious things in life
- The wonderful students over many years whose hard work and commitment always motivated us and put purpose in our writing
Welcome to the fifth edition of Basic Chemistry. This chemistry text was written and designed to prepare you for science-related professions, such as engineering, nursing, medicine, environmental or agricultural science, or for careers such as laboratory technology. This text assumes no prior knowledge of chemistry. Our main objective in writing this text is to make the study of chemistry an engaging and a positive experience for you by relating the structure and behavior of matter to real life. This new edition introduces more problem-solving strategies, more problem-solving guides, new Analyze the Problem with Connect features, new Try It First and Engage features, conceptual and challenge problems, and new sets of combined problems.

It is our goal to help you become a critical thinker by understanding scientific concepts that will form a basis for making important decisions about issues concerning health and the environment. Thus, we have utilized materials that

• help you to learn and enjoy chemistry
• relate chemistry to careers that interest you
• develop problem-solving skills that lead to your success in chemistry
• promote learning and success in chemistry

New for the Fifth Edition

New and updated features have been added throughout this fifth edition, including the following:

• NEW AND UPDATED! Chapter Openers provide timely examples and engaging, topical issues of the chemistry that is part of contemporary professions.
• NEW! A Follow Up story continues with material and questions related to the chapter opener.
• NEW! Engage feature asks students to think about the paragraph they are reading and to test their understanding by answering the Engage question in the margin, which is related to the topic.
• NEW! Try It First now precedes the Solution section of each Sample Problem to encourage the student to work on the problem before reading the given Solution.
• NEW! Connect feature added to Analyze the Problem boxes indicates the relationships between Given and Need.
• NEW! Applications are added to Questions and Problems sets that show the relevance between the chemistry content and the chapter opener story.
• NEW! A new topic with questions and problems on Hess’s Law, was added to Chapter 9.
• NEW! Interactive Videos give students the experience of step-by-step problem solving for problems from the text.
• UPDATED! Chapter Readiness sections at the beginning of each chapter list the Key Math Skills and Core Chemistry Skills from the previous chapters, which provide the foundation for learning new chemistry principles in the current chapter.
• UPDATED! Key Math Skills review basic math relevant to the chemistry the students are learning throughout the text. A Key Math Skill Review at the end of each chapter summarizes and gives additional examples.
• UPDATED! Core Chemistry Skills identify the key chemical principles in each chapter that are required for successfully learning chemistry. A Core Chemistry Skill Review at the end of each chapter helps reinforce the material and gives additional examples.
• UPDATED! Analyze the Problem features included in the Solutions of the Sample Problems strengthen critical-thinking skills and illustrate the breakdown of a word problem into the components required to solve it.
• UPDATED! Questions and Problems, Sample Problems, and art demonstrate the connection between the chemistry being discussed and how these skills will be needed in professional experience.
• UPDATED! Combining Ideas features offer sets of integrated problems that test students’ understanding by integrating topics from two or more previous chapters.

Chapter Organization of the Fifth Edition

In each textbook we write, we consider it essential to relate every chemical concept to real-life issues. Because a chemistry course may be taught in different time frames, it may be difficult to cover all the chapters in this text. However, each chapter is a complete package, which allows some chapters to be skipped or the order of presentation to be changed.

Chapter 1, Chemistry in Our Lives, discusses the Scientific Method in everyday terms, guides students in developing a study plan for learning chemistry, with a section of Key Math Skills that reviews the basic math, including scientific notation needed in chemistry calculations.

• The Chapter Opener and Follow Up feature the work and career of a forensic scientist.
• “Scientific Method: Thinking Like a Scientist” discusses the scientific method in everyday terms.
• A new Sample Problem requires the interpretation of a graph to determine the decrease in a child’s temperature when given Tylenol.
• Key Math Skills are: Identifying Place Values, Using Positive and Negative Numbers in Calculations including a new feature Calculator Operations, Calculating Percentages, Solving Equations, Interpreting Graphs, and Converting between Standard Numbers and Scientific Notation.

Chapter 2, Chemistry and Measurements, looks at measurement and emphasizes the need to understand numerical relationships of the metric system. Significant figures are discussed in the determination of final answers. Prefixes from the metric system are used to write equalities and conversion factors for problem-solving strategies. Density is discussed and used as a conversion factor.

• The Chapter Opener and Follow Up feature the work and career of a registered nurse.
• New photos, including an endoscope, a urine dipstick, a pint of blood, Keflex capsules, and salmon for omega-3 fatty acids, are added to improve visual introduction to clinical applications of chemistry.
• Updated Sample Problems relate questions and problem solving to health-related topics such as the measurements of blood volume, omega-3 fatty acids, radiological imaging, and medication orders.
• New Applications feature questions about measurements of daily values for minerals and vitamins, equalities and conversion factors for medications.
• A new Key Math Skill, Rounding Off, has been added.
• Core Chemistry Skills are: Counting Significant Figures, Using Significant Figures in Calculations, Using Prefixes, Writing Conversion Factors from Equalities, Using Conversion Factors, and Using Density as a Conversion Factor.

Chapter 3, Matter and Energy, classifies matter and states of matter, describes temperature measurement, and discusses energy, specific heat, and energy in nutrition. Physical and chemical changes and physical and chemical properties are discussed.

• The Chapter Opener and Follow Up feature the work and career of a dietitian.
• New Questions and Problems and Sample Problems include high temperatures used in cancer treatment, the energy produced by a high-energy shock output of a defibrillator, body temperature lowering using a cooling cap, and ice bag therapy for muscle injury.
• Core Chemistry Skills are: Classifying Matter, Identifying Physical and Chemical Changes, Converting between Temperature Scales, Using Energy Units, Calculating Specific Heat, and Using the Heat Equation.
• The interchapter problem set, Combining Ideas from Chapters 1 to 3, completes the chapter.

Chapter 4, Atoms and Elements, introduces elements and atoms and the periodic table element. The names and symbols of element 114, Flerovium, Fl, and element 116, Livermorium, Lv, are part of the periodic table. Atomic numbers and mass number are determined for isotopes. Atomic mass is calculated using the masses of the naturally occurring isotopes and their abundances.

• The Chapter Opener and Follow Up feature the work and career of a farmer.
• Atomic number and mass number are used to calculate the number of protons and neutrons in an atom.
• The number of protons and neutrons are used to calculate the mass number and to write the atomic symbol for an isotope.
• A weighted average analogy uses 8-lb and 14-lb bowling balls and the percent abundance of each to calculate weighted average of a bowling ball.
• Core Chemistry Skills are: Counting Protons and Neutrons, and Writing Atomic Symbols for Isotopes.

Chapter 5, Electronic Structure of Atoms and Periodic Trends, uses the electromagnetic spectrum to explain atomic spectra and develop the concept of energy levels and sublevels. Electrons in sublevels and orbitals are represented using orbital diagrams and electron configurations. Periodic properties of elements, including atomic radius and ionization energy, are related to their valence electrons. Small periodic tables illustrate the trends of periodic properties.

• The Chapter Opener and Follow Up feature the work and career of a materials engineer.
• The diagram for the electromagnetic spectrum has been updated.
• The three-dimensional representations of the s, p, and d orbitals are drawn.
• The trends in periodic properties are described for valence electrons, atomic size, ionization energy, and metallic character.
• Core Chemistry Skills are: Writing Electron Configurations, Using the Periodic Table to Write Electron Configurations, Identifying Trends in Periodic Properties, and Drawing Lewis Symbols.

Chapter 6, Ionic and Molecular Compounds, describes the formation of ionic and covalent bonds. Chemical formulas are written, and ionic compounds—including those with polyatomic ions—and molecular compounds are named.
• The Chapter Opener and Follow Up feature the work and career of a pharmacist.
• Core Chemistry Skills are: Writing Positive and Negative Ions, Writing Ionic Formulas, Naming Ionic Compounds, and Writing the Names and Formulas for Molecular Compounds.

Chapter 7, Chemical Quantities, discusses Avogadro’s number, the mole, and molar masses of compounds, which are used in calculations to determine the mass or number of particles in a quantity of a substance. The mass percent composition of a compound is calculated and used to determine its empirical and molecular formula.

• The Chapter Opener and Follow Up feature the work and career of a veterinarian.
• New and updated Guides to Problem Solving are: Converting the Moles (or Particles) of a Substance to Particles (or Moles), Calculating Moles of a Compound or Element, Calculating the Grams of an Element (or Compound) from the Grams of a Compound (or Element), and Calculating Mass Percent Composition from Molar Mass.
• Core Chemistry Skills are: Converting Particles to Moles, Calculating Molar Mass, Using Molar Mass as a Conversion Factor, Calculating Mass Percent Composition, Calculating an Empirical Formula, and Calculating a Molecular Formula.
• The interchapter problem set, Combining Ideas from Chapters 4 to 7, completes the chapter.

Chapter 8, Chemical Reactions introduces the method of balancing chemical equations, and discusses how to classify chemical reactions into types: combination, decomposition, single replacement, double replacement, and combustion reactions. A new section, Oxidation–Reduction Reactions, has been added.

• The Chapter Opener and Follow Up feature the work and career of an exercise physiologist.
• Core Chemistry Skills are: Balancing a Chemical Equation, Classifying Types of Chemical Reactions, and Identifying Oxidized and Reduced Substances.

Chapter 9, Chemical Quantities in Reactions, describes the mole and mass relationships among the reactants and products and provides calculations of limiting reactants and percent yields. A first section was divided into two new sections with an emphasis on the Law of Conservation of Mass.

• The Chapter Opener and Follow Up feature the work and career of an environmental scientist.
• Mole and mass relationships among the reactants and products are examined along with calculations of percent yield and limiting reactants.

• A new subsection, with questions and problems on Hess’s Law, was added.
• Core Chemistry Skills are: Using Mole–Mole Factors, Converting Grams to Grams, Calculating Quantity of Product from a Limiting Reactant, Calculating Percent Yield, and Using the Heat of Reaction.

Chapter 10, Properties of Solids and Liquids, introduces Lewis structures for molecules and ions with single and multiple bonds as well as resonance structures. Electronegativity leads to a discussion of the polarity of bonds and molecules. Lewis structures and VSEPR theory illustrate covalent bonding and the three-dimensional shapes of molecules and ions. The intermolecular forces between particles and their impact on states of matter and changes of state are described. The energy involved with changes of state is calculated.

• The Chapter Opener and Follow Up feature the work and career of a histologist.
• Lewis structures are drawn for molecules and ions with single, double, and triple bonds. Resonance structures are drawn if two or more Lewis structures are possible.
• Shapes and polarity of bonds and molecules are predicted using VSEPR theory.
• Intermolecular forces in compounds are discussed including ionic bonds, hydrogen bonds, dipole–dipole attractions, and dispersion forces.
• Core Chemistry Skills are Drawing Lewis Structures, Drawing Resonance Structures, Predicting Shape, Using Electronegativity, Identifying Polarity of Molecules, Identifying Intermolecular Forces, and Calculating Heat for Change of State.
• The interchapter problem set, Combining Ideas from Chapters 8 to 10, completes the chapter.

Chapter 11, Gases, discusses the properties of gases and calculates changes in gases using the gas laws: Boyle’s, Charles’s, Gay-Lussac’s, Avogadro’s, Dalton’s, and the Ideal Gas Law. Problem-solving strategies enhance the discussion and calculations with gas laws including chemical reactions using the ideal gas law.

• The Chapter Opener and Follow Up feature the work and career of a respiratory therapist.
• Applications includes calculations of mass or pressure of oxygen in uses of hyperbaric chambers.
• Core Chemistry Skills are: Using the Gas Laws, Using the Ideal Gas Law, Calculating Mass or Volume of a Gas in a Chemical Reaction, and Calculating Partial Pressure.

Chapter 12, Solutions, describes solutions, electrolytes, saturation and solubility, insoluble ionic compounds, concentrations, and osmosis. New problem-solving strategies clarify
the use of concentrations to determine volume or mass of solute. The volumes and concentrations of solutions are used in calculations of dilutions, reactions, and titrations. Properties of solutions, osmosis in the body, dialysis and changes in the freezing and boiling points of a solvent are discussed.

- The Chapter Opener and Follow Up feature the work and career of a dialysis nurse.
- Core Chemistry Skills are: Using Solubility Rules, Calculating Concentration, Using Concentration as a Conversion Factor, Calculating the Quantity of a Reactant or Product for a Chemical Reaction in Solution, and Calculating the Freezing Point/Boiling Point of a Solution.

Chapter 13, Reaction Rates and Chemical Equilibrium, looks at the rates of reactions and the equilibrium condition when forward and reverse rates for a reaction become equal. Equilibrium expressions for reactions are written and equilibrium constants are calculated. The equilibrium constant is used to calculate the concentration of a reactant or product at equilibrium. Le Châtelier’s principle is used to evaluate the impact on concentrations when stress is placed on a system at equilibrium. The concentrations of solutes in a solution is used to calculate the solubility product constant ($K_{sp}$).

- The Chapter Opener and Follow Up feature the work and career of a chemical oceanographer.
- New problems that visually represent equilibrium situations are added.
- Core Chemistry Skills are: Writing the Equilibrium Expression, Calculating an Equilibrium Constant, Calculating Equilibrium Concentrations, Using Le Châtelier’s Principle, Writing the Solubility Product Expression, Calculating a Solubility Product Constant, and Calculating the Molar Solubility.

Chapter 14, Acids and Bases, discusses acids and bases and their strengths, and conjugate acid–base pairs. The dissociation of strong and weak acids and bases is related to their strengths as acids or bases. The dissociation of water leads to the water dissociation expression, $K_w$, the pH scale, and the calculation of pH. Chemical equations for acids in reactions are balanced and titration of an acid is illustrated. Buffers are discussed along with their role in the blood. The pH of a buffer is calculated.

- The Chapter Opener and Follow Up feature work and career of a clinical laboratory technician.
- A new Guide to Writing the Acid Dissociation Expression has been added.
- Key Math Skills are: Calculating pH from $[H_3O^+]$, and Calculating $[H_3O^+]$ from pH.
- Core Chemistry Skills are: Identifying Conjugate Acid–Base Pairs, Calculating $[H_3O^+]$ and $[OH^-]$ in Solutions, Writing Equations for Reactions of Acids and Bases, Calculating Molarity or Volume of an Acid or Base in a Titration, and Calculating the pH of a Buffer.
- The interchapter problem set, Combining Ideas from Chapters 11 to 14, completes the chapter.

Chapter 15, Oxidation and Reduction, looks at the characteristics of oxidation and reduction reactions. Oxidation numbers are assigned to the atoms in elements, molecules, and ions to determine the components that lose electrons during oxidation and gain electrons during reduction. The half-reaction method is utilized to balance oxidation–reduction reactions. The production of electrical energy in voltaic cells and the requirement of electrical energy in electrolytic cells are diagrammed using half-cells. The activity series is used to determine the spontaneous direction of an oxidation–reduction reaction.

- The Chapter Opener and Follow Up feature the work and career of a dentist.
- A new Guide to Identifying Oxidizing and Reducing Agents has been added.
- Core Chemistry Skills are: Assigning Oxidation Numbers, Using Oxidation Numbers, Identifying Oxidizing and Reducing Agents, Using Half-Reactions to Balance Redox Equations, and Identifying Spontaneous Reactions.

Chapter 16, Nuclear Chemistry, looks at the types of radiation emitted from the nuclei of radioactive atoms. Nuclear equations are written and balanced for both naturally occurring radioactivity and artificially produced radioactivity. The half-lives of radioisotopes are discussed, and the amount of time for a sample to decay is calculated. Radioisotopes important in the field of nuclear medicine are described. Fission and fusion and their role in energy production are discussed.

- The Chapter Opener and Follow Up feature the work and career of a radiologist.
- Core Chemistry Skills are: Writing Nuclear Equations, and Using Half-Lives.
- The interchapter problem set, Combining Ideas from Chapters 15 and 16, completes the chapter.

Chapter 17, Organic Chemistry, compares inorganic and organic compounds, and describes the condensed and line-angle structural formulas of alkanes, alkenes, alcohols, ethers, aldehydes, ketones, carboxylic acids, esters, amines, and amides.

- The Chapter Opener and Follow Up feature the work and career of a firefighter/emergency medical technician.
- The properties of organic and inorganic compounds are now compared in Table 17.1.
Acknowledgments

The preparation of a new text is a continuous effort of many people. As in our work on other textbooks, we are thankful for the support, encouragement, and dedication of many people who put in hours of tireless effort to produce a high-quality book that provides an outstanding learning package. The editorial team at Pearson Publishing has done an exceptional job. We want to thank, Jeanne Zalesky, editor in chief, and Editors Terry Haugen and Scott Dustan, who supported our vision of this fifth edition and the development of new problem-solving strategies.

We much appreciate all the wonderful work of project manager Laura Perry, who was like an angel encouraging us at each step, while skillfully coordinating reviews, art, web site materials, and all the things it takes to make a book come together. We appreciate the work of Lisa Pierce, program manager, and Lindsay Bethoney of Lumina Datamatics, Inc., who brilliantly coordinated all phases of the manuscript to the final pages of a beautiful book. Thanks to Mark Quirie, manuscript and accuracy reviewer, and copy editors of Lumina Datamatics, Inc., who precisely analyzed and edited the initial and final manuscripts and pages to make sure the words and problems were correct to help students learn chemistry. Their keen eyes and thoughtful comments were extremely helpful in the development of this text.

We are especially proud of the art program in this text, which lends beauty and understanding to chemistry. We would like to thank Marilyn Perry and Gary Hespenheide, interior and cover design managers and book designer, whose creative ideas provided the outstanding design for the cover and pages of the book. Erica Gordon, photo researcher, was invaluable in researching and selecting vivid photos for the text so that students can see the beauty of chemistry. Thanks also to Bio-Rad Laboratories for their courtesy and use of Know-ItAll ChemWindows drawing software that helped us produce chemical structures for the manuscript. The macro-to-micro illustrations designed by Production Solutions and Precision Graphics give students visual impressions of the atomic and molecular organization of everyday things and are a fantastic learning tool. We also appreciate all the hard work put in by the marketing team in the field and Executive Marketing Manager, Chris Barker.
We are extremely grateful to an incredible group of peers for their careful assessment of all the new ideas for the text; for their suggested additions, corrections, changes, and deletions; and for providing an incredible amount of feedback about improvements for the book.

If you would like to share your experience with chemistry, or have questions and comments about this text, we would appreciate hearing from you.

Karen and Bill Timberlake
Email: khemist@aol.com

FAVORITE QUOTES

The whole art of teaching is only the art of awakening the natural curiosity of young minds.
—Anatole France

One must learn by doing the thing; though you think you know it, you have no certainty until you try.
—Sophocles

Discovery consists of seeing what everybody has seen and thinking what nobody has thought.
—Albert Szent-Györgyi

I never teach my pupils; I only attempt to provide the conditions in which they can learn.
—Albert Einstein

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Accuracy Reviewer

Mark Quirie  
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The EPA has set the maximum level for mercury in tuna at 0.5 ppm.

There are 0.5 mg of propranolol in 1 tablet.

It has one significant figure. The 0.5 mg is measured:

The 40. mg is measured:

Follow Up

GREG'S VISIT WITH HIS DOCTOR
On Greg's last visit to his doctor, he complained of feeling tired. His doctor orders a blood test for iron. Sandra, the registered nurse, draws a venipuncture and withdraws 80 mL of blood. About 70% of the iron in the body is used to form hemoglobin, which is a protein in the red blood cells that...
### Volume (V)

The volume of gas equals the size of the container in which the gas is placed. When you inflate a tire on a basketball, you are adding some gas particles. The increase in the number of particles hitting the walls of the tire or basketball increases the volume. Sometimes, on a cold morning, a tire looks flat. The volume of the tire has decreased because a lower temperature decreases the speed of the molecules, which in turn reduces the force of their impacts on the walls of the tire. The most common units for volume measurement are liters (L) and milliliters (mL).

### TRY IT FIRST

In a polar covalent bond, the electrons share electrons unequally. For example, 1 mol of carbon contains 6 electrons. Alkynes Melting occurs when the particles have sufficient kinetic energy and rate of motion to overcome the forces holding them together. Amides

- **System of Units** (SI)  
- **Gases** and the atom with the fewest decimal places.

### KEY MATH SKILL

Calculating pH from \([\text{H}_2\text{O}^+]\)

### TOPIC REVIEW 17.1 Alkanes

**LEARNING GOAL** Write the IUPAC names and draw the condensed or line-angle structural formulas for alkanes.

### KEY TERMS

- **Avogadro’s number** The number of items in a mole, equal to \(6.02 \times 10^{23}\).
- **empirical formula** The simplest or smallest whole-number ratio of the atoms in a formula.
- **formula unit** The group of ions represented by the formula of an ionic compound.

### CHAPTER REVIEW

**LEARNING GOAL** Describe the intermolecular forces between ions, polar covalent molecules, and nonpolar covalent molecules.

### KEY TERMS with definitions are listed at the end of each chapter as well as in the Glossary/Index at the end of the text.

### Concept Maps at the end of each chapter show how all the key concepts fit together.

### TRY IT FIRST

Timberlake’s accessible **Writing Style** is based on careful development of chemical concepts suited to the skills and backgrounds of students in chemistry. The TRY IT FIRST feature encourages you to try to solve the problem before you compare your work with the Solution.

### Benefit

- **Help you focus your studying by emphasizing what is most important in each section.**
- **Helps you understand new terms and chemical concepts.**
- **Help you master the basic quantitative skills to succeed in chemistry.**
- **Help you master the basic problem-solving skills needed to succeed in chemistry.**
- **Helps you identify what you know about the solution and what you need to learn.**
- **Help you determine your mastery of the chapter concepts and study for your tests.**
- **Help you recall the important new terms in each chapter.**
- **Encourage learning by giving you a visual guide to the interrelationship among all the concepts new to each chapter.**

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- **Help you recall the important new terms in each chapter.**
- **Encourage learning by giving you a visual guide to the interrelationship among all the concepts new to each chapter.**
How many significant figures are in each of the following?

- 19.7616
- 2

Identify the exact number (s), if any, in each of the following:

a. 3 hamburgers and 6 oz of hamburger

b. 

c. 

d. 

Determine if the bonds are polar covalent or nonpolar covalent.

b.

Sample Problems illustrate worked-out solutions with explanations and required calculations. Study Checks associated with each Sample Problem allow you to check your problem-solving strategies with the Answer.

Understanding the Concepts are questions with visual representations placed at the end of each chapter.

Additional Questions and Problems at the end of each chapter provide further study and application of the topics from the entire chapter. Problems are paired and the Answers to the odd-numbered problems are given at the end of each chapter.

Challenge Questions at the end of each chapter provide complex questions. Answers to the odd-numbered questions are given at the end of each chapter.

Combining Ideas are sets of integrated problems placed after every two to four chapters that are useful as practice exams. Answers to the odd-numbered problems are given at the end of each Combining Idea.

Rounding Off

- To round to the nearest whole number, drop the digits at the end of the number.
- To round to one decimal place, drop all digits at the end of the number after the first decimal place.
- To round to two decimal places, drop all digits at the end of the number after the second decimal place.
### Resources

*Basic Chemistry, fifth edition, provides an integrated teaching and learning package of support material for both students and professors.*

<table>
<thead>
<tr>
<th>Name of Supplement</th>
<th>Available in Print</th>
<th>Available Online</th>
<th>Instructor or Student Supplement</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>MasteringChemistry® (<a href="http://www.masteringchemistry.com">www.masteringchemistry.com</a>)</td>
<td>✓</td>
<td>Resource for Students and Instructors</td>
<td>MasteringChemistry® from Pearson is the leading online teaching and learning system designed to improve results by engaging students before, during, and after class with powerful content. Ensure that students arrive ready to learn by assigning educationally effective content before class, and encourage critical thinking and retention with in-class resources such as Learning Catalytics. Students can further master concepts after class through traditional homework assignments that provide hints and answer-specific feedback. The Mastering gradebook records scores for all automatically graded assignments while diagnostic tools give instructors access to rich data to assess student understanding and misconceptions.</td>
<td></td>
</tr>
<tr>
<td>Pearson eText</td>
<td>✓</td>
<td>Resource for Students</td>
<td>The fifth edition of <em>Basic Chemistry</em> features a Pearson eText enhanced with media within Mastering. In conjunction with Mastering assessment capabilities, Videos, and animations will improve student engagement and knowledge retention. Additionally, the Pearson eText offers students the power to create notes, highlight text in different colors, create bookmarks, zoom, and view single or multiple pages.</td>
<td></td>
</tr>
<tr>
<td>Instructor’s Solutions Manual–Download Only</td>
<td>✓</td>
<td>Resource for Instructors</td>
<td>Prepared by Mark Quirie, the solutions manual highlights chapter topics, and includes answers and solutions for all questions and problems in the text.</td>
<td></td>
</tr>
<tr>
<td>Instructor Resource Materials–Download Only</td>
<td>✓</td>
<td>Resource for Instructors</td>
<td>Includes all the art, photos, and tables from the book in JPEG format for use in classroom projection or when creating study materials and tests. In addition, the instructors can access modifiable PowerPoint™ lecture outlines. Also available are downloadable files of the Instructor’s Solutions Manual and a set of “clicker questions” designed for use with classroom-response systems. Also visit the Pearson Education catalog page for Timberlake’s <em>Basic Chemistry</em>, fifth edition, at <a href="http://www.pearsonglobaleditions.com/timberlake">www.pearsonglobaleditions.com/timberlake</a> to download available instructor supplements.</td>
<td></td>
</tr>
<tr>
<td>TestGen Test Bank–Download Only</td>
<td>✓</td>
<td>Resource for Instructors</td>
<td>Prepared by William Timberlake, this resource includes more than 2000 questions in multiple-choice, matching, true/false, and short-answer format.</td>
<td></td>
</tr>
<tr>
<td>Laboratory Manual by Karen Timberlake</td>
<td>✓</td>
<td>Resource for Students</td>
<td>This best-selling lab manual coordinates 35 experiments with the topics in <em>Basic Chemistry</em>, fifth edition, uses laboratory investigations to explore chemical concepts, develop skills of manipulating equipment, reporting data, solving problems, making calculations, and drawing conclusions.</td>
<td></td>
</tr>
</tbody>
</table>
Follow Ups and Applications

Chapter Openers throughout the text connect chemistry to real life. Each chapter begins with an image and details of a profession such as engineering, medicine, environmental science or agriculture, or exercise physiology. Follow Ups at the end of chapter discuss the chemistry in the Chapter Opener and include Applications. These questions show students how the chemistry they are learning applies specifically to their professional careers.

Follow Up
FORENSIC EVIDENCE SOLVES THE MURDER

Using a variety of laboratory tests, Sarah finds ethylene glycol in the victim’s blood. The quantitative tests indicate that the victim had ingested 125 g of ethylene glycol. Sarah determines that the liquid in a glass found at the crime scene was ethylene glycol that had been added to an alcoholic beverage. Ethylene glycol is a clear, sweet-tasting, thick liquid that is odorous and miscible with water. It is easy to confuse since it is used as antifreeze in automobiles and in brake fluid. Because the initial symptoms of ethylene glycol poisoning are similar to being intoxicated, the victim’s relatives assume it was present.

An expert in forensic chemistry, Pat, can save deposition of the scene (evidence): Systemic vascular damage, and kidney failure. A chemical analysis (hemoglobin) may be used to remove ethylene glycol from the blood. A toxic amount of ethylene glycol is 1.5 g of ethylene glycol per 1000 g of body mass. Thus, 75 g would be lethal for a 50 kg (110 lb) person.

Mark determines that fingerprints on the glass containing the ethylene glycol were those of the victim’s husband. The evidence along with the container of antifreeze found in the victim’s home led to the arrest and conviction of the husband for poisoning his wife.

Key Math Skills - Writing numbers in scientific notation

To write a number in scientific notation, the number is written as a product of a number greater than or equal to 1 but less than 10 and a power of 10. The power of 10 is given by the number of places to move the decimal point. The decimal point is moved to obtain a coefficient that is at least 1 but less than 10.

1.5 Writing numbers in scientific notation

a. 0.000 016
b. 3.5

Focusing on New Problem-Solving Strategies

This new edition introduces more problem-solving strategies, more problem-solving guides, new Analyze the Problem with Connect features, new Try It First and Engage features, conceptual and challenge problems, and new sets of combined problems.

NEW! Connect feature has been added to the Analyze the Problem boxes, which specifies the information that relates the Given and Need sections.

NEW! Try It First now precedes the Solution section of each Sample Problem to encourage the student to work on the problem before reading the given Solution.

NEW! Engage feature asks students to think about the paragraph they are reading and to test their understanding by answering the Engage question in the margin, which is related to the topic.

Sample Problem 1.7 Scientific Notation

Write each of the following in scientific notation:

a. 3.500
b. 0.000 016

What is different and what is the same for an atom of Sn-105 and an atom of Sn-132?
Interactive videos and demonstrations help students through some of the more challenging topics by showing how chemistry works in real life and introducing a bit of humor into chemical problem solving and demonstrations. Topics include Using Conversion Factors, Balancing Nuclear Equations, and Chemical vs. Physical Change.

Sample Calculations walk students through the most challenging chemistry problems and provide a fresh perspective on how to approach individual problems and plan solutions. Topics include Using Conversion Factors, Mass Calculations for Reactions, and Concentration of Solutions.

Green play button icons appear in the margins throughout the text. In the eText, the icons link to new interactive videos that the student can use to clarify and reinforce important concepts. All Interactive Videos are available in web and mobile-friendly formats through the eText, and are assignable activities in MasteringChemistry.

**Interactive Videos**

> **SAMPLE PROBLEM 1.5 Solving Equations**
> Solve the following equation for \( V_2 \):
> 
> \[ P_1 V_1 = P_2 V_2 \]

**TRY IT FIRST**

**SOLUTION**

\[ \frac{P_1 V_1}{P_1} = \frac{P_2 V_2}{P_1} \]

To solve for \( V_2 \), divide both sides by the symbol \( P_1 \).

\[ \frac{V_1}{P_1} = \frac{P_2}{P_1} \]

\[ V_2 = \frac{P_2}{P_1} \]

**STUDY CHECK 1.5**
MasteringChemistry® from Pearson is the leading online teaching and learning system designed to improve results by engaging students before, during, and after class with powerful content. Instructors may ensure that students arrive ready to learn by assigning educationally effective content before class, and encourage critical thinking and retention with in-class resources such as Learning Catalytics. Students can further master concepts after class through traditional homework assignments that provide hints and answer-specific feedback. The Mastering gradebook records scores for all automatically graded assignments while diagnostic tools give instructors access to rich data to assess student understanding and misconceptions.

Mastering brings learning full circle by continuously adapting to each student and making learning more personal than ever—before, during, and after class.

### Before Class

#### Dynamic Study Modules

Help students quickly learn chemistry! Now assignable, Dynamic Study Modules (DSMs) enable your students to study on their own and be better prepared with the basic math and chemistry skills needed to succeed in the course. The mobile app is available for iOS and Android devices for study on the go and results can be tracked in the MasteringChemistry gradebook.

#### Reading Quizzes

Reading Quizzes give instructors the opportunity to assign reading and test students on their comprehension of chapter content.

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28
During Class

Learning Catalytics

Learning Catalytics is a “bring your own device” student engagement, assessment, and classroom intelligence system. With Learning Catalytics you can:

- Assess students in real time, using open-ended tasks to probe student understanding.
- Understand immediately where students are and adjust your lecture accordingly.
- Manage student interactions with intelligent grouping and timing.

After Class

Tutorials and Coaching

Students learn chemistry by practicing chemistry. Tutorials, featuring specific wrong-answer feedback, hints, and a wide variety of educationally effective content, guide your students through the toughest topics in Basic Chemistry.
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A CALL CAME IN TO 911 from a man who found his wife lying on the floor of their home. When the police arrived, they determined that the woman was dead. The husband said he had worked late, and just arrived home. The victim’s body was lying on the floor of the living room. There was no blood at the scene, but the police did find a glass on the side table that contained a small amount of liquid. In an adjacent laundry room/garage, the police found a half-empty bottle of antifreeze. The bottle, glass, and liquid were bagged and sent to the forensic laboratory.

In another 911 call, a man was found lying on the grass outside his home. Blood was present on his body, and some bullet casings were found on the grass. Inside the victim’s home, a weapon was recovered. The bullet casings and the weapon were bagged and sent to the forensic laboratory.

Sarah and Mark, forensic scientists, use scientific procedures and chemical tests to examine the evidence from law enforcement agencies. Sarah proceeds to analyze blood, stomach contents, and the unknown liquid from the first victim’s home. She will look for the presence of drugs, poisons, and alcohol. Her lab partner Mark will analyze the fingerprints on the glass. He will also match the characteristics of the bullet casings to the weapon that was found at the second crime scene.

CAREER

Forensic Scientist

Most forensic scientists work in crime laboratories that are part of city or county legal systems where they analyze bodily fluids and tissue samples collected by crime scene investigators. In analyzing these samples, forensic scientists identify the presence or absence of specific chemicals within the body to help solve the criminal case. Some of the chemicals they look for include alcohol, illegal or prescription drugs, poisons, arson debris, metals, and various gases such as carbon monoxide. In order to identify these substances, a variety of chemical instruments and highly specific methodologies are used. Forensic scientists also analyze samples from criminal suspects, athletes, and potential employees. They also work on cases involving environmental contamination and animal samples for wildlife crimes. Forensic scientists usually have a bachelor’s degree that includes courses in math, chemistry, and biology.
1.1 Chemistry and Chemicals

LEARNING GOAL Define the term chemistry and identify substances as chemicals.

Now that you are in a chemistry class, you may be wondering what you will be learning. What questions in science have you been curious about? Perhaps you are interested in how smog is formed or how aspirin relieves a headache. Just like you, chemists are curious about the world we live in.

How does car exhaust produce the smog that hangs over our cities? One component of car exhaust is nitrogen oxide (NO), which forms in car engines where high temperatures convert nitrogen gas ($N_2$) and oxygen gas ($O_2$) to NO. In the atmosphere, the NO($g$) reacts with O$_2$(g) to form NO$_2$(g), which has a reddish brown color of smog. In chemistry, reactions are written in the form of equations:

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

Smog

Why does aspirin relieve a headache? When a part of the body is injured, substances called prostaglandins are produced, which cause inflammation and pain. Aspirin acts to block the production of prostaglandins, thereby reducing inflammation, pain, and fever. Chemists in the medical field develop new treatments for diabetes, genetic defects, cancer, AIDS, and other diseases. Chemists in the environmental field study the ways in which human development impacts the environment and develop processes that help reduce environmental degradation. For the chemist in the forensic laboratory, the nurse in the dialysis unit, the dietitian, the chemical engineer, or the agricultural scientist, chemistry plays a central role in understanding problems, assessing possible solutions, and making important decisions.

Chemistry

Chemistry is the study of the composition, structure, properties, and reactions of matter. Matter is another word for all the substances that make up our world. Perhaps you imagine that chemistry takes place only in a laboratory where a chemist is working in a white coat and goggles. Actually, chemistry happens all around you every day and has an impact on everything you use and do. You are doing chemistry when you cook food, add bleach to your laundry, or start your car. A chemical reaction has taken place when silver tarnishes or an antacid tablet fizzes when dropped into water. Plants grow because chemical reactions convert carbon dioxide, water, and energy to carbohydrates. Chemical reactions take place when you digest food and break it down into substances that you need for energy and health.

Chemicals

A chemical is a substance that always has the same composition and properties wherever it is found. All the things you see around you are composed of one or more chemicals. Chemical processes take place in chemistry laboratories, manufacturing plants, and pharmaceutical labs as well as every day in nature and in our bodies. Often the terms chemical and substance are used interchangeably to describe a specific type of matter.
Every day, you use products containing substances that were developed and prepared by chemists. Soaps and shampoos contain chemicals that remove oils on your skin and scalp. When you brush your teeth, the substances in toothpaste clean your teeth, prevent plaque formation, and stop tooth decay. Some of the chemicals used to make toothpaste are listed in TABLE 1.1.

In cosmetics and lotions, chemicals are used to moisturize, prevent deterioration of the product, fight bacteria, and thicken the product. Your clothes may be made of natural materials, such as cotton, or synthetic substances, such as nylon or polyester. Perhaps you wear a ring or watch made of gold, silver, or platinum. Your breakfast cereal is probably fortified with iron, calcium, and phosphorus, whereas the milk you drink is enriched with vitamins A and D. Antioxidants are chemicals added to food to prevent it from spoiling. Some of the chemicals you may encounter when you cook in the kitchen are shown in FIGURE 1.1.

**TABLE 1.1** Chemicals Commonly Used in Toothpaste

<table>
<thead>
<tr>
<th>Chemical</th>
<th>Function</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calcium carbonate</td>
<td>Used as an abrasive to remove plaque</td>
</tr>
<tr>
<td>Sorbitol</td>
<td>Prevents loss of water and hardening of toothpaste</td>
</tr>
<tr>
<td>Sodium lauryl sulfate</td>
<td>Used to loosen plaque</td>
</tr>
<tr>
<td>Titanium dioxide</td>
<td>Makes toothpaste white and opaque</td>
</tr>
<tr>
<td>Triclosan</td>
<td>Inhibits bacteria that cause plaque and gum disease</td>
</tr>
<tr>
<td>Sodium fluorophosphate</td>
<td>Prevents formation of cavities by strengthening tooth enamel with fluoride</td>
</tr>
<tr>
<td>Methyl salicylate</td>
<td>Gives toothpaste a pleasant wintergreen flavor</td>
</tr>
</tbody>
</table>

Silicon dioxide (glass)

Metal alloy

Chemically treated water

Natural polymers

Natural gas

Fruits grown with fertilizers and pesticides

**FIGURE 1.1** Many of the items found in a kitchen are chemicals or products of chemical reactions.

What are some other chemicals found in a kitchen?

**Branches of Chemistry**

The field of chemistry is divided into several branches. General chemistry is the study of the composition, properties, and reactions of matter. Organic chemistry is the study of substances that contain the element carbon. Biological chemistry is the study of the chemical reactions that take place in biological systems. Today chemistry is often combined with other sciences, such as geology and physics, to form cross-disciplines such as geochemistry and physical chemistry. Geochemistry is the study of the chemical composition of ores, soils, and minerals of the surface of the Earth and other planets. Physical chemistry is the study of the physical nature of chemical systems, including energy changes.
When you were very young, you explored the things around you by touching and tasting. As you grew, you asked questions about the world in which you live. What is lightning? Where does a rainbow come from? Why is water blue? As an adult, you may have wondered how antibiotics work or why vitamins are important to your health. Every day, you ask questions and seek answers to organize and make sense of the world around you.

When the late Nobel Laureate Linus Pauling described his student life in Oregon, he recalled that he read many books on chemistry, mineralogy, and physics. “I mulled over the properties of materials: why are some substances colored and others not, why are some minerals or inorganic compounds hard and others soft?” He said, “I was building up this tremendous background of empirical knowledge and at the same time asking a great number of questions.” Linus Pauling won two Nobel Prizes: the first, in 1954, was in chemistry for his work on the nature of chemical bonds and the determination of the structures of complex substances; the second, in 1962, was the Peace Prize.

The Scientific Method

The process of trying to understand nature is unique to each scientist. However, the scientific method is a process that scientists use to make observations in nature, gather data, and explain natural phenomena.

1. **Observations** The first step in the scientific method is to make observations about nature and ask questions about what you observe. When an observation always seems to be true, it may be stated as a law that predicts that behavior and is often measurable. However, a law does not explain that observation. For example, we can use the Law of Gravity to predict that if we drop our chemistry book it would fall on the table or the floor but this law does not explain why our book falls.

2. **Hypothesis** A scientist forms a hypothesis, which gives a possible explanation of an observation or a law. The hypothesis must be stated in such a way that it can be tested by experiments.

3. **Experiments** To determine if a hypothesis is true or false, experiments are done to find a relationship between the hypothesis and the observations. The results of the experiments may confirm the hypothesis. However, if the experiments do not confirm the hypothesis, it is modified or discarded. Then new experiments will be designed to test the hypothesis.

4. **Conclusion/Theory** When the results of the experiments are analyzed, a conclusion is made as to whether the hypothesis is true or false. When experiments give consistent results, the hypothesis may be stated to be true. Even then, the hypothesis
Using the Scientific Method in Everyday Life

You may be surprised to realize that you use the scientific method in your everyday life. Suppose you visit a friend in her home. Soon after you arrive, your eyes start to itch and you begin to sneeze. Then you observe that your friend has a new cat. Perhaps you ask yourself why you are sneezing and you form the hypothesis that you are allergic to cats. To test your hypothesis, you leave your friend’s home. If the sneezing stops, perhaps your hypothesis is correct. You test your hypothesis further by visiting another friend who also has a cat. If you start to sneeze again, your experimental results support your hypothesis and you come to the conclusion that you are allergic to cats. However, if you continue sneezing after you leave your friend’s home, your hypothesis is not supported. Now you need to form a new hypothesis, which could be that you have a cold.

**SAMPLE PROBLEM 1.1 Scientific Method**

Identify each of the following statements as an observation (O), a hypothesis (H), or an experiment (E):

a. A silver tray turns a dull gray color when left uncovered.

b. When a silver tray is covered with plastic wrap, it does not tarnish.

c. Oxygen reacts with silver when the tray is exposed to air.
TRY IT FIRST
SOLUTION
a. observation (O)  b. experiment (E)  c. hypothesis (H)

STUDY CHECK 1.1
The following statements are found in a student’s notebook. Identify each of the following as an observation (O), a hypothesis (H), or an experiment (E):

a. “Today I placed two tomato seedlings in the garden, and two more in a closet. I will give all the plants the same amount of water and fertilizer.”
b. “After 50 days, the tomato plants in the garden are 3 ft high with green leaves. The plants in the closet are 8 in. tall and yellow.”
c. “Tomato plants need sunlight to grow.”

ANSWER
a. experiment (E)  b. observation (O)  c. hypothesis (H)

QUESTIONS AND PROBLEMS

1.2 Scientific Method: Thinking Like a Scientist

LEARNING GOAL Describe the activities that are part of the scientific method.

1.7 Define each of the following terms of the scientific method:
   a. hypothesis  b. experiment  c. theory  d. observation

1.8 Identify each of the following activities in the scientific method as an observation (O), a hypothesis (H), an experiment (E), or a conclusion (C):
   a. Formulate a possible explanation for your experimental results.
   b. Make notes about nature.
   c. Design an experimental plan that will give new information about a problem.
   d. State a generalized summary of your experimental results.

Applications

1.9 Identify each activity, a to f, as an observation (O), a hypothesis (H), an experiment (E), or a conclusion (C).
   At a popular restaurant, where Chang is the head chef, the following occurred:
   a. Chang determined that sales of the house salad had dropped.
   b. Chang decided that the house salad needed a new dressing.
   c. In a taste test, Chang prepared four bowls of lettuce, each with a new dressing: sesame seed, olive oil and balsamic vinegar, creamy Italian, and blue cheese.
   d. The tasters rated the sesame seed salad dressing as the favorite.
   e. After two weeks, Chang noted that the orders for the house salad with the new sesame seed dressing had doubled.
   f. Chang decided that the sesame seed dressing improved the sales of the house salad because the sesame seed dressing enhanced the taste.

1.10 Identify each activity, a to f, as an observation (O), a hypothesis (H), an experiment (E), or a conclusion (C).
   Lucia wants to develop a process for dyeing shirts so that the color will not fade when the shirt is washed. She proceeds with the following activities:
   a. Lucia notices that the dye in a design fades when the shirt is washed.
   b. Lucia decides that the dye needs something to help it combine with the fabric.
   c. She places a spot of dye on each of four shirts and then places each one separately in water, salt water, vinegar, and baking soda and water.
   d. After one hour, all the shirts are removed and washed with a detergent.
   e. Lucia notices that the dye has faded on the shirts in water, salt water, and baking soda, whereas the dye did not fade on the shirt soaked in vinegar.
   f. Lucia thinks that the vinegar binds with the dye so it does not fade when the shirt is washed.

1.11 Identify each of the following as an observation (O), a hypothesis (H), an experiment (E), or a conclusion (C):
   a. One hour after drinking a glass of regular milk, Jim experienced stomach cramps.
   b. Jim thinks he may be lactose intolerant.
   c. Jim drinks a glass of lactose-free milk and does not have any stomach cramps.
   d. Jim drinks a glass of regular milk to which he has added lactase, an enzyme that breaks down lactose, and has no stomach cramps.

1.12 Identify each of the following as an observation (O), a hypothesis (H), an experiment (E), or a conclusion (C):
   a. Sally thinks she may be allergic to shrimp.
   b. Yesterday, one hour after Sally ate a shrimp salad, she broke out in hives.
   c. Today, Sally had some soup that contained shrimp, but she did not break out in hives.
   d. Sally realizes that she does not have an allergy to shrimp.
1.3 Learning Chemistry: A Study Plan

**LEARNING GOAL** Develop a study plan for learning chemistry.

Here you are taking chemistry, perhaps for the first time. Whatever your reasons for choosing to study chemistry, you can look forward to learning many new and exciting ideas.

**Features in This Text Help You Study Chemistry**

This text has been designed with study features to complement your individual learning style. On the inside of the front cover is a periodic table of the elements. On the inside of the back cover are tables that summarize useful information needed throughout your study of chemistry. Each chapter begins with Looking Ahead, which outlines the topics in the chapter. Key Terms are bolded when they first appear in the text, and are summarized at the end of each chapter. They are also listed and defined in the comprehensive Glossary and Index, which appears at the end of the text. Key Math Skills and Core Chemistry Skills that are critical to learning chemistry are indicated by icons in the margin, and summarized at the end of each chapter. In the Chapter Readiness list at the beginning of every chapter, the Key Math Skills and Core Chemistry Skills from previous chapters related to the current chapter concepts are highlighted for your review.

Before you begin reading, obtain an overview of a chapter by reviewing the topics in Looking Ahead. As you prepare to read a section of the chapter, look at the section title and turn it into a question. For example, for section 1.1, “Chemistry and Chemicals,” you could ask, “What is chemistry?” or “What are chemicals?” When you come to a Sample Problem, take the time to work it through and compare your solution to the one provided. As you read the text, you will see Engage features in the margin, which remind you to pause your reading and interact with a question related to the material.

The Try It First feature above the Solution of each Sample Problem is a reminder for you to work out the problem before you look at the Solution. Many Sample Problems are accompanied by a Guide to Problem Solving, which gives the steps needed to work the problem. The Analyze the Problem feature in some Sample Problems includes Given, the information we have; Need, what we are going to accomplish; and Connect, how we proceed from Given to Need. When you compare your answer with the Solution provided, you know what you need to correct or change. This process of trying the problem first will help you develop successful problem solving techniques. Then work the associated Study Check. The answers to all the Study Checks are included and you can compare your answer to the one provided.

At the end of each chapter section, you will find a set of Questions and Problems that allows you to apply problem solving immediately to the new concepts. The problems are paired, which means that each of the odd-numbered problems is matched to the following even-numbered problem. At the end of each chapter, the answers to all the odd-numbered problems are provided. If the answers match yours, you most likely understand the topic; if not, you need to study the section again.

Throughout each chapter, boxes titled “Chemistry Link to Health” and “Chemistry Link to the Environment” help you connect the chemical concepts you are learning to real-life situations. Many of the figures and diagrams use macro-to-micro illustrations to depict the atomic level of organization of ordinary objects, such as the atoms in aluminum foil. These visual models illustrate the concepts described in the text and allow you to “see” the world in a microscopic way.

At the end of each chapter, you will find several study aids that complete the chapter. Chapter Reviews provide a summary in easy-to-read bullet points and Concept Maps visually show the connections between important topics. The Key Terms, which are in boldface type within the chapter, are listed with their definitions. Understanding the Concepts, a set of questions that use art and models, helps you visualize concepts. Additional Questions and Problems and Challenge Problems provide additional exercises to test your understanding of the topics in the chapter. Applications are groups of problems that apply section content to current topics. Answers to all of the odd-numbered problems complete the chapter and you can compare your answers to the ones provided.

After some chapters, problem sets called Combining Ideas test your ability to solve problems containing material from more than one chapter.
Using Active Learning

A student who is an active learner continually interacts with the chemical ideas while reading the text, working problems, and attending lectures. Let’s see how this is done.

As you read and practice problem solving, you remain actively involved in studying, which enhances the learning process. In this way, you learn a small amount of information and establish the necessary foundation for understanding the next section. You may also note questions you have about the reading, which you can discuss with your professor or laboratory instructor. TABLE 1.2 summarizes these steps for active learning. The time you spend in a lecture is a useful learning time. By keeping track of the class schedule and reading the assigned material before a lecture, you become aware of the new terms and concepts you need to learn. Some questions that occur during your reading may be answered during the lecture. If not, you can ask your professor for further clarification.

<table>
<thead>
<tr>
<th>TABLE 1.2  Steps in Active Learning</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Read each Learning Goal for an overview of the material.</td>
</tr>
<tr>
<td>2. Form a question from the title of the section you are going to read.</td>
</tr>
<tr>
<td>3. Read the section, looking for answers to your question.</td>
</tr>
<tr>
<td>4. Self-test by working Sample Problems and Study Checks.</td>
</tr>
<tr>
<td>5. Complete the Questions and Problems that follow that section, and check the answers for the magenta odd-numbered problems at the end of the chapter.</td>
</tr>
<tr>
<td>6. Proceed to the next section and repeat the steps.</td>
</tr>
</tbody>
</table>

Many students find that studying with a group can be beneficial to learning. In a group, students motivate each other to study, fill in gaps, and correct misunderstandings by teaching and learning together. Studying alone does not allow the process of peer correction. In a group, you can cover the ideas more thoroughly as you discuss the reading and problem solve with other students. You may find that it is easier to retain new material and new ideas if you study in short sessions throughout the week rather than all at once. Waiting to study until the night before an exam does not give you time to understand concepts and practice problem solving.

Making a Study Plan

As you embark on your journey into the world of chemistry, think about your approach to studying and learning chemistry. You might consider some of the ideas in the following list. Check those ideas that will help you successfully learn chemistry. Commit to them now. Your success depends on you.

My study plan for learning chemistry will include the following:

- reading the chapter before lecture
- going to lecture
- reviewing the Learning Goals
- keeping a problem notebook
- reading the text as an active learner
- answering the Engage questions
- trying to work the Sample Problem before looking at the Solution
- working the Questions and Problems at the end of each section and checking answers
SAMPLE PROBLEM 1.2 A Study Plan for Learning Chemistry

Which of the following activities would you include in your study plan for learning chemistry successfully?

a. skipping lecture
b. going to the professor’s office hours
c. keeping a problem notebook
d. waiting to study until the night before the exam
e. trying to work the Sample Problem before looking at the Solution

TRY IT FIRST

SOLUTION

Your success in chemistry can be improved by:

b. going to the professor’s office hours
c. keeping a problem notebook
e. trying to work the Sample Problem before looking at the Solution

STUDY CHECK 1.2

Which of the following will help you learn chemistry?

a. skipping review sessions
b. working assigned problems
c. staying up all night before an exam
d. reading the assignment before a lecture

ANSWER

b and d

QUESTIONS AND PROBLEMS

1.3 Learning Chemistry: A Study Plan

LEARNING GOAL Develop a study plan for learning chemistry.

1.13 What are four things you can do to help yourself to succeed in chemistry?

1.14 What are four things that would make it difficult for you to learn chemistry?

1.15 A student in your class asks you for advice on learning chemistry. Which of the following might you suggest?

a. forming a study group
b. skipping a lecture
c. visiting the professor during office hours
d. waiting until the night before an exam to study
e. answering the Engage question

1.16 A student in your class asks you for advice on learning chemistry. Which of the following might you suggest?

a. doing the assigned problems
b. not reading the text; it’s never on the test
c. attending review sessions
d. reading the assignment before a lecture
e. keeping a problem notebook
1.4 Key Math Skills for Chemistry

LEARNING GOAL Review math concepts used in chemistry: place values, positive and negative numbers, percentages, solving equations, and interpreting graphs.

During your study of chemistry, you will work many problems that involve numbers. You will need various math skills and operations. We will review some of the key math skills that are particularly important for chemistry. As we move through the chapters, we will also reference the key math skills as they apply.

Identifying Place Values

For any number, we can identify the place value for each of the digits in that number. These place values have names such as the ones place (first place to the left of the decimal point) or the tens place (second place to the left of the decimal point). A premature baby has a mass of 2518 g. We can indicate the place values for the number 2518 as follows:

<table>
<thead>
<tr>
<th>Digit</th>
<th>Place Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>thousands</td>
</tr>
<tr>
<td>5</td>
<td>hundreds</td>
</tr>
<tr>
<td>1</td>
<td>tens</td>
</tr>
<tr>
<td>8</td>
<td>ones</td>
</tr>
</tbody>
</table>

We also identify place values such as the tenths place (first place to the right of the decimal point) and the hundredths place (second place to the right of the decimal place). A silver coin has a mass of 6.407 g. We can indicate the place values for the number 6.407 as follows:

<table>
<thead>
<tr>
<th>Digit</th>
<th>Place Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>6</td>
<td>ones</td>
</tr>
<tr>
<td>4</td>
<td>tenths</td>
</tr>
<tr>
<td>0</td>
<td>hundredths</td>
</tr>
<tr>
<td>7</td>
<td>thousandths</td>
</tr>
</tbody>
</table>

Note that place values ending with the suffix *ths* refer to the decimal places to the right of the decimal point.

SAMPLE PROBLEM 1.3 Identifying Place Values

A bullet found at a crime scene has a mass of 15.24 g. What are the place values for the digits in the mass of the bullet?

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>Digit</th>
<th>Place Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>tens</td>
</tr>
<tr>
<td>5</td>
<td>ones</td>
</tr>
<tr>
<td>2</td>
<td>tenths</td>
</tr>
<tr>
<td>4</td>
<td>hundredths</td>
</tr>
</tbody>
</table>

STUDY CHECK 1.3

A bullet found at a crime scene contains 0.925 g of lead. What are the place values for the digits in the mass of the lead?
Using Positive and Negative Numbers in Calculations

A positive number is any number that is greater than zero and has a positive sign (+). Often the positive sign is understood and not written in front of the number. For example, the number +8 can also be written as 8. A negative number is any number that is less than zero and is written with a negative sign (−). For example, a negative eight is written as −8.

Multiplication and Division of Positive and Negative Numbers

When two positive numbers or two negative numbers are multiplied, the answer is positive (+).

\[2 \times 3 = +6\]
\[(-2) \times (-3) = +6\]

When a positive number and a negative number are multiplied, the answer is negative (−).

\[2 \times (-3) = -6\]
\[(-2) \times 3 = -6\]

The rules for the division of positive and negative numbers are the same as the rules for multiplication. When two positive numbers or two negative numbers are divided, the answer is positive (+).

\[\frac{6}{3} = 2\]
\[\frac{-6}{-3} = 2\]

When a positive number and a negative number are divided, the answer is negative (−).

\[\frac{-6}{3} = -2\]
\[\frac{6}{-3} = -2\]

Addition of Positive and Negative Numbers

When positive numbers are added, the sign of the answer is positive.

\[3 + 4 = 7\]

The + sign (+7) is understood.

When negative numbers are added, the sign of the answer is negative.

\[(-3) + (-4) = -7\]

When a positive number and a negative number are added, the smaller number is subtracted from the larger number, and the result has the same sign as the larger number.

\[12 + (-15) = -3\]

Subtraction of Positive and Negative Numbers

When two numbers are subtracted, change the sign of the number to be subtracted and follow the rules for addition shown above.

\[12 - (+5) = 12 - 5 = 7\]
\[12 - (-5) = 12 + 5 = 17\]
\[-12 - (-5) = -12 + 5 = -7\]
\[-12 - (+5) = -12 - 5 = -17\]
Calculator Operations
On your calculator, there are four keys that are used for basic mathematical operations. The change sign \( \pm \) key is used to change the sign of a number.

To practice these basic calculations on the calculator, work through the problem going from the left to the right doing the operations in the order they occur. If your calculator has a change sign \( \pm \) key, a negative number is entered by pressing the number and then pressing the change sign \( \pm \) key. At the end, press the equals \( = \) key or ANS or ENTER.

Addition and Subtraction

| Example 1: \( 15 - 8 + 2 = \) | Solution: \( 15 \text{ } 8 \text{ } + \text{ } 2 \text{ } = \text{ } 9 \) |
| Example 2: \( 4 + (-10) - 5 = \) | Solution: \( 4 \text{ } + \text{ } 10 \text{ } \pm \text{ } 5 \text{ } = \text{ } -11 \) |

Multiplication and Division

| Example 3: \( 2 \times (-3) = \) | Solution: \( 2 \text{ } \times \text{ } 3 \text{ } \pm \text{ } 6 \) |
| Example 4: \( \frac{8 \times 3}{4} = \) | Solution: \( 8 \text{ } 	ext{ } \times 	ext{ } 3 \text{ } \div \text{ } 4 \text{ } = \text{ } 6 \) |

Calculating Percentages

To determine a percentage, divide the parts by the total (whole) and multiply by 100%. For example, if an aspirin tablet contains 325 mg of aspirin (active ingredient) and the tablet has a mass of 545 mg, what is the percentage of aspirin in the tablet?

\[
\frac{325 \text{ mg aspirin}}{545 \text{ mg tablet}} \times 100\% = 59.6\% \text{ aspirin}
\]

When a value is described as a percentage (%), it represents the number of parts of an item in 100 of those items. If the percentage of red balls is 5, it means there are 5 red balls in every 100 balls. If the percentage of green balls is 50, there are 50 green balls in every 100 balls.

\[
5\% \text{ red balls} = \frac{5 \text{ red balls}}{100 \text{ balls}} \quad 50\% \text{ green balls} = \frac{50 \text{ green balls}}{100 \text{ balls}}
\]

SAMPLE PROBLEM 1.4 Calculating a Percentage

A bullet found at a crime scene may be used as evidence in a trial if the percentage of metals is a match to the composition of metals in a bullet from the suspect’s ammunition. If a bullet found at a crime scene contains 13.9 g of lead, 0.3 g of tin, and 0.9 g of antimony, what is the percentage of each metal in the bullet? Express your answers to the ones place.
SOLUTION

Total mass = 13.9 g + 0.3 g + 0.9 g = 15.1 g

Percentage of lead
\[ \frac{13.9 \text{ g}}{15.1 \text{ g}} \times 100\% = 92\% \text{ lead} \]

Percentage of tin
\[ \frac{0.3 \text{ g}}{15.1 \text{ g}} \times 100\% = 2\% \text{ tin} \]

Percentage of antimony
\[ \frac{0.9 \text{ g}}{15.1 \text{ g}} \times 100\% = 6\% \text{ antimony} \]

STUDY CHECK 1.4

A bullet seized from the suspect’s ammunition has a composition of lead 11.6 g, tin 0.5 g, and antimony 0.4 g.

a. What is the percentage of each metal in the bullet? Express your answers to the ones place.

b. Could the bullet removed from the suspect’s ammunition be considered as evidence that the suspect was at the crime scene mentioned in Sample Problem 1.4?

ANSWER

a. The bullet from the suspect’s ammunition is lead 93%, tin 4%, and antimony 3%.
b. The composition of this bullet does not match the bullet from the crime scene and cannot be used as evidence.

Solving Equations

In chemistry, we use equations that express the relationship between certain variables. Let’s look at how we would solve for \( x \) in the following equation:

\[ 2x + 8 = 14 \]

Our overall goal is to rearrange the items in the equation to obtain \( x \) on one side.

1. Place all like terms on one side. The numbers 8 and 14 are like terms. To remove the 8 from the left side of the equation, we subtract 8. To keep a balance, we need to subtract 8 from the 14 on the other side.

\[ 2x + 8 - 8 = 14 - 8 \]
\[ 2x = 6 \]

2. Isolate the variable you need to solve for. In this problem, we obtain \( x \) by dividing both sides of the equation by 2. The value of \( x \) is the result when 6 is divided by 2.

\[ \frac{2x}{2} = \frac{6}{2} \]
\[ x = 3 \]

3. Check your answer. Check your answer by substituting your value for \( x \) back into the original equation.

\[ 2(3) + 8 = 14 \]
\[ 6 + 8 = 14 \]
\[ 14 = 14 \]  Your answer \( x = 3 \) is correct.

Summary: To solve an equation for a particular variable, be sure you perform the same mathematical operations on both sides of the equation.

- If you eliminate a symbol or number by subtracting, you need to subtract that same symbol or number on the opposite side.
- If you eliminate a symbol or number by adding, you need to add that same symbol or number on the opposite side.
If you cancel a symbol or number by dividing, you need to divide both sides by that same symbol or number.

If you cancel a symbol or number by multiplying, you need to multiply both sides by that same symbol or number.

When we work with temperature, we may need to convert between degrees Celsius and degrees Fahrenheit using the following equation:

\[ T_F = 1.8(T_C) + 32 \]

To obtain the equation for converting degrees Fahrenheit to degrees Celsius, we subtract 32 from both sides.

\[ T_F - 32 = 1.8(T_C) + 32 - 32 \]

\[ T_F - 32 = 1.8(T_C) \]

To obtain \( T_C \) by itself, we divide both sides by 1.8.

\[ \frac{T_F - 32}{1.8} = \frac{1.8(T_C)}{1.8} = T_C \]

### SAMPLE PROBLEM 1.5 Solving Equations

Solve the following equation for \( V_2 \):

\[ P_1V_1 = P_2V_2 \]

**TRY IT FIRST**

**SOLUTION**

\[ P_1V_1 = P_2V_2 \]

To solve for \( V_2 \), divide both sides by the symbol \( P_2 \).

\[ \frac{P_1V_1}{P_2} = \frac{P_2V_2}{P_2} \]

\[ V_2 = \frac{P_1V_1}{P_2} \]

**STUDY CHECK 1.5**

Solve the following equation for \( m \):

\[ \text{heat} = m \times \Delta T \times SH \]

**ANSWER**

\[ m = \frac{\text{heat}}{\Delta T \times SH} \]

### Interpreting Graphs

A graph represents the relationship between two variables. These quantities are plotted along two perpendicular axes, which are the \( x \) axis (horizontal) and \( y \) axis (vertical).

**Example**

In the graph Volume of a Balloon Versus Temperature, the volume of a gas in a balloon is plotted against its temperature.

**Title**

Look at the title. What does it tell us about the graph? The title indicates that the volume of a balloon was measured at different temperatures.
Vertical Axis
Look at the label and the numbers on the vertical (y) axis. The label indicates that the volume of the balloon was measured in liters (L). The numbers, which are chosen to include the low and high measurements of the volume of the gas, are evenly spaced from 22.0 L to 30.0 L.

Horizontal Axis
The label on the horizontal (x) axis indicates that the temperature of the balloon was measured in degrees Celsius (°C). The numbers are measurements of the Celsius temperature, which are evenly spaced from 0 °C to 100 °C.

Points on the Graph
Each point on the graph represents a volume in liters that was measured at a specific temperature. When these points are connected, a line is obtained.

Interpreting the Graph
From the graph, we see that the volume of the gas increases as the temperature of the gas increases. This is called a direct relationship. Now we use the graph to determine the volume at various temperatures. For example, suppose we want to know the volume of the gas at 50 °C. We would start by finding 50 °C on the x axis and then drawing a line up to the plotted line. From there, we would draw a horizontal line that intersects the y axis and read the volume value where the line crosses the y axis as shown on the graph above.

**SAMPLE PROBLEM 1.6 Interpreting a Graph**
A nurse administers Tylenol to lower a child’s fever. The graph shows the body temperature of the child plotted against time.

a. What is measured on the vertical axis?

b. What is the range of values on the vertical axis?

c. What is measured on the horizontal axis?

d. What is the range of values on the horizontal axis?

**TRY IT FIRST**

**SOLUTION**

a. body temperature, in degrees Celsius

b. 37.0 °C to 39.4 °C

c. time, in minutes, after Tylenol was given

d. 0 min to 30 min
STUDY CHECK 1.6

a. Using the graph in Sample Problem 1.6, what was the child’s temperature 15 min after Tylenol was given?
b. How many minutes elapsed before the temperature decreased to 38.0 °C?

ANSWER

a. 37.6 °C  
b. 8 min

QUESTIONS AND PROBLEMS

1.4 Key Math Skills for Chemistry

LEARNING GOAL Review math concepts used in chemistry:
place values, positive and negative numbers, percentages, solving equations, and interpreting graphs.

1.17 What is the place value for the bold digit?
   a. 7.3288
   b. 16.1234
   c. 4675.99

1.18 What is the place value for the bold digit?
   a. 97.5689
   b. 375.88
   c. 46.1000

1.19 Evaluate each of the following:
   a. 15 − (−8) =
   b. −8 + (−22) =
   c. 4 × (−2) + 6 =

1.20 Evaluate each of the following:
   a. −11 − (−9) =
   b. 34 + (−55) =
   c. \( \frac{-56}{8} \) =

Applications

1.21 a. A clinic had 25 patients on Friday morning. If 21 patients were given flu shots, what percentage of the patients received flu shots? Express your answer to the ones place.
b. An alloy contains 56 g of pure silver and 22 g of pure copper. What is the percentage of silver in the alloy? Express your answer to the ones place.
c. A collection of coins contains 11 nickels, 5 quarters, and 7 dimes. What is the percentage of dimes in the collection? Express your answer to the ones place.

1.22 a. At a local hospital, 35 babies were born. If 22 were boys, what percentage of the newborns were boys? Express your answer to the ones place.
b. An alloy contains 67 g of pure gold and 35 g of pure zinc. What is the percentage of zinc in the alloy? Express your answer to the ones place.
c. A collection of coins contains 15 pennies, 14 dimes, and 6 quarters. What is the percentage of pennies in the collection? Express your answer to the ones place.

1.23 Solve each of the following for \( a \):
   a. \( 4a + 4 = 40 \)  
b. \( \frac{a}{6} = 7 \)

1.24 Solve each of the following for \( b \):
   a. \( 2b + 7 = b + 10 \)  
b. \( 3b - 4 = 24 - b \)

Use the following graph for questions 1.25 and 1.26:

1.25 a. What does the title indicate about the graph?
b. What is measured on the vertical axis?
c. What is the range of values on the vertical axis?
d. Does the temperature increase or decrease with an increase in time?

1.26 a. What is measured on the horizontal axis?
b. What is the range of values on the horizontal axis?
c. What is the temperature of the tea after 20 min?
d. How many minutes were needed to reach a temperature of 45 °C?
1.5 Writing Numbers in Scientific Notation

LEARNING GOAL Write a standard number in scientific notation and vice versa.

In chemistry, we use numbers that are very large and very small. We might measure something as tiny as the width of a human hair, which is about 0.000 008 m. Or perhaps we want to count the number of hairs on the average human scalp, which is about 100 000 hairs. In this text, we add spaces between sets of three digits when it helps make the places easier to count. However, we will see that it is more convenient to write large and small numbers in scientific notation.

A number written in scientific notation has two parts: a coefficient and a power of 10. For example, the number 2400 is written in scientific notation as $2.4 \times 10^3$. The coefficient, 2.4, is obtained by moving the decimal point to the left to give a number that is at least 1 but less than 10. Because we moved the decimal point three places to the left, the power of 10 is a positive 3, which is written as $10^3$. When a number greater than 1 is converted to scientific notation, the power of 10 is positive.

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### Standard Number | Scientific Notation
---|---
$2 400$ | $2.4 \times 10^3$

In another example, 0.000 86 is written in scientific notation as $8.6 \times 10^{-4}$. The coefficient, 8.6, is obtained by moving the decimal point to the right. Because the decimal point is moved four places to the right, the power of 10 is a negative 4, written as $10^{-4}$. When a number less than 1 is written in scientific notation, the power of 10 is negative.

### Standard Number | Scientific Notation
---|---
0.000 86 | $8.6 \times 10^{-4}$

**TABLE 1.3** gives some examples of numbers written as positive and negative powers of 10. The powers of 10 are a way of keeping track of the decimal point in the number. **TABLE 1.4** gives several examples of writing measurements in scientific notation.
Scientific Notation and Calculators

You can enter a number in scientific notation on many calculators using the [EE or EXP] key. After you enter the coefficient, press the [EE or EXP] key and enter only the power of 10, because the [EE or EXP] key already includes the \( \times 10 \) value. To enter a negative power of 10, press the [+] or the [-] key, depending on your calculator.

<table>
<thead>
<tr>
<th>Number to Enter</th>
<th>Procedure</th>
<th>Calculator Display</th>
</tr>
</thead>
<tbody>
<tr>
<td>( 4 \times 10^6 )</td>
<td>( 4 \ [\text{EE or EXP}] \ 6 )</td>
<td>( 4.06 ) or ( 4E06 )</td>
</tr>
<tr>
<td>( 2.5 \times 10^{-4} )</td>
<td>( 2.5 \ [\text{EE or EXP}] \ [+/-] \ 4 )</td>
<td>( 2.5\text{E}04 ) or ( 2.5\text{E}04 )</td>
</tr>
</tbody>
</table>

When a calculator display appears in scientific notation, it is shown as a number that is at least 1 but less than 10, followed by a space and the power of 10. To express this display in scientific notation, write the coefficient value, write \( \times 10 \), and use the power of 10 as an exponent.
1.5 Writing Numbers in Scientific Notation

On many scientific calculators, a number is converted into scientific notation using the appropriate keys. For example, the number 0.000 52 is entered, followed by pressing the 2nd or 3rd function key and the SCI key. The scientific notation appears in the calculator display as a coefficient and the power of 10.

\[ 0.000 \, 52 \, \text{SCI} = \, 5.2 \times 10^{-4} \]

**SAMPLE PROBLEM 1.7 Scientific Notation**

Write each of the following in scientific notation:

a. 3500  
b. 0.000 016

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th><strong>ANALYZE THE PROBLEM</strong></th>
<th><strong>Given</strong></th>
<th><strong>Need</strong></th>
<th><strong>Connect</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>standard number</td>
<td>scientific notation</td>
<td>coefficient is at least 1 but less than 10</td>
<td></td>
</tr>
</tbody>
</table>

**a. 3500**

**STEP 1** Move the decimal point to obtain a coefficient that is at least 1 but less than 10. For a number greater than 1, the decimal point is moved to the left three places to give a coefficient of 3.5.

**STEP 2** Express the number of places moved as a power of 10. Moving the decimal point three places to the left gives a power of 3, written as $10^3$.

**STEP 3** Write the product of the coefficient multiplied by the power of 10.

\[ 3.5 \times 10^3 \]

**b. 0.000 016**

**STEP 1** Move the decimal point to obtain a coefficient that is at least 1 but less than 10. For a number less than 1, the decimal point is moved to the right five places to give a coefficient of 1.6.

**STEP 2** Express the number of places moved as a power of 10. Moving the decimal point five places to the right gives a power of negative 5, written as $10^{-5}$.

**STEP 3** Write the product of the coefficient multiplied by the power of 10.

\[ 1.6 \times 10^{-5} \]

**STUDY CHECK 1.7**

Write each of the following in scientific notation:

a. 425 000  
b. 0.000 000 86

**ANSWER**

a. $4.25 \times 10^5$  
b. $8.6 \times 10^{-7}$

---

<table>
<thead>
<tr>
<th><strong>Calculator Display</strong></th>
<th><strong>Expressed in Scientific Notation</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>7.52 04 or 7.52×10^4 or 7.52E04</td>
<td>7.52 × 10^4</td>
</tr>
<tr>
<td>5.8−02 or 5.8×10^−2 or 5.8E−02</td>
<td>5.8 × 10^−2</td>
</tr>
</tbody>
</table>
Converting Scientific Notation to a Standard Number

When a number written in scientific notation has a positive power of 10, the standard number is obtained by moving the decimal point to the right for the same number of places as the power of 10. Placeholder zeros are used, as needed, to give additional places.

<table>
<thead>
<tr>
<th>Scientific Notation</th>
<th>Standard Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>(4.3 \times 10^2)</td>
<td>(430)</td>
</tr>
</tbody>
</table>

For a number with a negative power of 10, the standard number is obtained by moving the decimal point to the left for the same number of places as the power of 10. Placeholder zeros are added in front of the coefficient as needed.

<table>
<thead>
<tr>
<th>Scientific Notation</th>
<th>Standard Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>(2.5 \times 10^{-5})</td>
<td>(0.000025)</td>
</tr>
</tbody>
</table>

SAMPLE PROBLEM 1.8 Writing Scientific Notation as a Standard Number

Write each of the following as a standard number:

a. \(7.2 \times 10^{-3}\)  
   TRY IT FIRST
   SOLUTION
   a. To write the standard number for an exponential number with a negative power of 10, move the decimal point to the left the same number of places (three) as the power of 10. Add placeholder zeros before the coefficient as needed.
   \(7.2 \times 10^{-3} = 0.0072\) m
   b. To write the standard number for an exponential number with a positive power of 10, move the decimal point to the right the same number of places (five) as the power of 10. Add placeholder zeros following the coefficient as needed.
   \(2.4 \times 10^5 = 240,000\)

STUDY CHECK 1.8
Write \(7.25 \times 10^{-4}\) as a standard number.

ANSWER
0.000 725

QUESTIONS AND PROBLEMS
1.5 Writing Numbers in Scientific Notation

LEARNING GOAL: Write a standard number in scientific notation and vice versa.

1.27 Write each of the following in scientific notation:
   a. 55 000
   b. 480
   c. 0.000 005
   d. 0.001 4
   e. 0.0072

1.28 Write each of the following in scientific notation:
   a. 180 000 000
   b. 0.000 006
   c. 750
   d. 0.15
   e. 0.024

1.29 Write each of the following as a standard number:
   a. \(1.2 \times 10^4\)
   b. \(8.25 \times 10^{-2}\)
   c. \(4 \times 10^6\)
   d. \(5.8 \times 10^{-3}\)

1.30 Write each of the following as a standard number:
   a. \(3.6 \times 10^{-3}\)
   b. \(8.75 \times 10^7\)
   c. \(3 \times 10^{-2}\)
   d. \(2.12 \times 10^1\)

1.31 Which number in each of the following pairs is larger?
   a. \(7.2 \times 10^3\) or \(8.2 \times 10^2\)
   b. \(4.5 \times 10^{-4}\) or \(3.2 \times 10^{-2}\)
   c. \(1 \times 10^4\) or \(1 \times 10^{-4}\)
   d. \(0.000 52\) or \(6.8 \times 10^{-2}\)

1.32 Which number in each of the following pairs is smaller?
   a. \(4.9 \times 10^{-3}\) or \(5.5 \times 10^{-9}\)
   b. \(1250\) or \(3.4 \times 10^2\)
   c. \(0.000 004\) or \(5.0 \times 10^2\)
   d. \(2.5 \times 10^2\) or \(4 \times 10^3\)
Follow Up

**FORENSIC EVIDENCE SOLVES THE MURDER**

Using a variety of laboratory tests, Sarah finds ethylene glycol in the victim’s blood. The quantitative tests indicate that the victim had ingested 125 g of ethylene glycol. Sarah determines that the liquid in a glass found at the crime scene was ethylene glycol that had been added to an alcoholic beverage. Ethylene glycol is a clear, sweet-tasting, thick liquid that is odorless and mixes with water. It is easy to obtain since it is used as antifreeze in automobiles and in brake fluid. Because the initial symptoms of ethylene glycol poisoning are similar to being intoxicated, the victim is often unaware of its presence.

If ingestion of ethylene glycol occurs, it can cause depression of the central nervous system, cardiovascular damage, and kidney failure. If discovered quickly, hemodialysis may be used to remove ethylene glycol from the blood. A toxic amount of ethylene glycol is 1.5 g of ethylene glycol/kg of body mass. Thus, 75 g could be fatal for a 50-kg (110 lb) person.

Mark determines that fingerprints on the glass containing the ethylene glycol were those of the victim’s husband. This evidence along with the container of antifreeze found in the home led to the arrest and conviction of the husband for poisoning his wife.

**Applications**

1.33 A container was found in the home of the victim that contained 120 g of ethylene glycol in 450 g of liquid. What was the percentage of ethylene glycol? Express your answer to the ones place.

1.34 If the toxic quantity is 1.5 g of ethylene glycol per 1000 g of body mass, what percentage of ethylene glycol is fatal?
1.1 Chemistry and Chemicals

**LEARNING GOAL** Define the term chemistry and identify substances as chemicals.

- Chemistry is the study of the composition, structure, properties, and reactions of matter.
- A chemical is any substance that always has the same composition and properties wherever it is found.

1.2 Scientific Method: Thinking Like a Scientist

**LEARNING GOAL** Describe the activities that are part of the scientific method.

- The scientific method is a process of explaining natural phenomena beginning with making observations, forming a hypothesis, and performing experiments.
- After repeated successful experiments, a hypothesis may become a theory.

1.3 Learning Chemistry: A Study Plan

**LEARNING GOAL** Develop a study plan for learning chemistry.

- A study plan for learning chemistry utilizes the features in the text and develops an active learning approach to study.

### KEY TERMS

- **chemical**: A substance that has the same composition and properties wherever it is found.
- **chemistry**: The study of the composition, structure, properties, and reactions of matter.
- **conclusion**: An explanation of an observation that has been validated by repeated experiments that support a hypothesis.
- **experiment**: A procedure that tests the validity of a hypothesis.
- **hypothesis**: An unverified explanation of a natural phenomenon.
- **observation**: Information determined by noting and recording a natural phenomenon.

### KEY MATH SKILLS

The chapter section containing each Key Math Skill is shown in parentheses at the end of each heading.

#### Identifying Place Values (1.4)

- The place value identifies the numerical value of each digit in a number.

**Example**: Identify the place value for each of the digits in the number 456.78.

**Answer**: | Digit | Place Value |
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>4</td>
<td>hundreds</td>
</tr>
<tr>
<td>5</td>
<td>tens</td>
</tr>
<tr>
<td>6</td>
<td>ones</td>
</tr>
<tr>
<td>7</td>
<td>tenths</td>
</tr>
<tr>
<td>8</td>
<td>hundredths</td>
</tr>
</tbody>
</table>
Using Positive and Negative Numbers in Calculations (1.4)

- A positive number is any number that is greater than zero and has a positive sign (+). A negative number is any number that is less than zero and is written with a negative sign (−).
- When two positive numbers are added, multiplied, or divided, the answer is positive.
- When two negative numbers are multiplied or divided, the answer is positive. When two negative numbers are added, the answer is negative.
- When a positive and a negative number are multiplied or divided, the answer is negative.
- When two positive numbers are added, multiplied, or divided, the answer is positive. When two negative numbers are added, the answer is negative.
- When two negative numbers are multiplied or divided, the answer is positive.
- When a positive and a negative number are added, the smaller number is subtracted from the larger number and the result has the same sign as the larger number.
- When two numbers are subtracted, change the sign of the number to be subtracted then follow the rules for addition.

Example: Evaluate each of the following:

a. \(-8 - 14 = _____\)
b. \(6 \times (-3) = _____\)

Answer: a. \(-8 - 14 = -22\)
b. \(6 \times (-3) = -18\)

Calculating Percentages (1.4)

- A percentage is the part divided by the total (whole) multiplied by 100%.

Example: A drawer contains 6 white socks and 18 black socks. What is the percentage of white socks?

Answer: \(\frac{6 \text{ white socks}}{24 \text{ total socks}} \times 100\% = 25\% \text{ white socks}\)

Solving Equations (1.4)

An equation in chemistry often contains an unknown. To rearrange an equation to obtain the unknown factor by itself, you keep it balanced by performing matching mathematical operations on both sides of the equation.

- If you eliminate a number or symbol by subtracting, subtract that same number or symbol on the opposite side.
- If you eliminate a number or symbol by adding, add that same number or symbol on the opposite side.
- If you cancel a number or symbol by dividing, divide both sides by that same number or symbol.
- If you cancel a number or symbol by multiplying, multiply both sides by that same number or symbol.

Example: Solve the equation for \(a\):

\[3a - 8 = 28\]

Answer: Add 8 to both sides: \[3a - 8 + 8 = 28 + 8\]

\[3a = 36\]

Divide both sides by 3:

\[\frac{3a}{3} = \frac{36}{3}\]

\[a = 12\]

Check:

\[3(12) - 8 = 28\]

\[36 - 8 = 28\]

\[28 = 28\]

Your answer \(a = 12\) is correct.

Interpreting Graphs (1.4)

- A graph represents the relationship between two variables.
- The quantities are plotted along two perpendicular axes, which are the \(x\) axis (horizontal) and \(y\) axis (vertical).
- The title indicates the components of the \(x\) and \(y\) axes.
- Numbers on the \(x\) and \(y\) axes show the range of values of the variables.
- The graph shows the relationship between the component on the \(y\) axis and that on the \(x\) axis.

Example:

![Graph of Solubility of Sugar in Water Versus Temperature](image)

- a. Does the amount of sugar that dissolves in 100 mL of water increase or decrease when the temperature increases?
- b. How many grams of sugar dissolve in 100 mL of water at 70 °C?
- c. At what temperature (°C) will 275 g of sugar dissolve in 100 mL of water?

Answer: a. increase  
   b. 320 g  
   c. 55 °C

Converting between Standard Numbers and Scientific Notation (1.5)

- A number written in scientific notation consists of a coefficient and a power of 10.

A number is written in scientific notation by:

- Moving the decimal point to obtain a coefficient that is at least 1 but less than 10.
- Expressing the number of places moved as a power of 10. The power of 10 is positive if the decimal point is moved to the left, negative if the decimal point is moved to the right.
- The standard number is obtained by moving the decimal point for the same number of places as the power of 10.

Example: Write the number 28 000 in scientific notation.

Answer: Moving the decimal point four places to the left gives a coefficient of 2.8 and a positive power of 10, \(10^4\). The number 28 000 written in scientific notation is \(2.8 \times 10^4\).

Example: Write \(5.6 \times 10^{-5}\) as a standard number.

Answer: Moving the decimal point five places to the left and adding placeholder zeros needed gives 0.000 056.
UNDERSTANDING THE CONCEPTS

The chapter sections to review are shown in parentheses at the end of each question.

1.35 A “chemical-free” shampoo includes the following ingredients: water, cocamide, glycerin, and citric acid. Is the shampoo truly “chemical-free”? (1.1)

1.36 A “chemical-free” sunscreen includes the following ingredients: titanium dioxide, vitamin E, and vitamin C. Is the sunscreen truly “chemical-free”? (1.1)

1.37 According to Sherlock Holmes, “One must follow the rules of scientific inquiry, gathering, observing, and testing data, then formulating, modifying, and rejecting hypotheses, until only one remains.” Did Holmes use the scientific method? Why or why not? (1.2)

1.38 In A Scandal in Bohemia, Sherlock Holmes receives a mysterious note. He states, “I have no data yet. It is a capital mistake to theorize before one has data. Insensibly one begins to twist facts to suit theories, instead of theories to suit facts.” What do you think Holmes meant? (1.2)

ADDitional QUESTIONS AND PROBLEMS

1.43 Identify each of the following as an observation (O), a hypothesis (H), an experiment (E), or a conclusion (C): (1.2)
   a. During an assessment in the emergency room, a nurse writes that the patient has a resting pulse of 30 beats/min.
   b. A nurse thinks that an incision from a recent surgery that is red and swollen is infected.
   c. Repeated studies show that lowering sodium in the diet leads to a decrease in blood pressure.

1.44 Identify each of the following as an observation (O), a hypothesis (H), an experiment (E), or a conclusion (C): (1.2)
   a. Drinking coffee at night keeps me awake.
   b. If I stop drinking coffee in the afternoon, I will be able to sleep at night.
   c. I will try drinking coffee only in the morning.

1.45 Select the correct phrase(s) to complete the following statement: If experimental results do not support your hypothesis, you should: (1.2)
   a. pretend that the experimental results support your hypothesis
   b. modify your hypothesis
   c. do more experiments

1.39 Classify each of the following statements as an observation (O) or a hypothesis (H): (1.2)
   a. You are likely to get food poisoning from expired food products.
   b. Adding 5 g of salt to ice decreases the melting temperature to −1 °C.
   c. Adding salt to ice can decrease the melting temperature.

1.40 Classify each of the following statements as an observation (O) or a hypothesis (H): (1.2)
   a. Analysis of a sample of 100 toys indicated that 5% contained lead.
   b. Reduced CO₂ emission can help reduce global warming.
   c. A child with fever higher than 40 °C is likely to suffer damage to brain function.

1.41 For each of the following, indicate if the answer has a positive or negative sign: (1.4)
   a. Two negative numbers are multiplied.
   b. A larger positive number is added to a smaller negative number.

1.42 For each of the following, indicate if the answer has a positive or negative sign: (1.4)
   a. A negative number is divided by a positive number.
   b. Two negative numbers are added.

1.46 Select the correct phrase(s) to complete the following statement: A hypothesis is confirmed when: (1.2)
   a. one experiment proves the hypothesis
   b. many experiments validate the hypothesis
   c. you think your hypothesis is correct

1.47 Which of the following will help you develop a successful study plan? (1.3)
   a. skipping lecture and just reading the text
   b. working the Sample Problems as you go through a chapter
   c. going to your professor’s office hours
   d. reading through the chapter, but working the problems later

1.48 Which of the following will help you develop a successful study plan? (1.3)
   a. studying all night before the exam
   b. forming a study group and discussing the problems together
   c. working problems in a notebook for easy reference
   d. copying the homework answers from a friend

1.49 Evaluate each of the following: (1.4)
   a. 5 × (−0.5) =
   b. −1023 − 854 =
   c. \(-\frac{273}{-3}\) =

1.50 Evaluate each of the following: (1.4)
   a. \(-24−(−56)\) =
   b. \(\frac{562}{-8}\) =
   c. 32 × (−0.6) =
1.51 A bag of gumdrops contains 16 orange gumdrops, 8 yellow gumdrops, and 16 black gumdrops. (1.4)
   a. What is the percentage of yellow gumdrops? Express your answer to the ones place.
   b. What is the percentage of black gumdrops? Express your answer to the ones place.

1.52 On the first chemistry test, 12 students got As, 18 students got Bs, and 20 students got Cs. (1.4)
   a. What is the percentage of students who received Bs? Express your answer to the ones place.
   b. What is the percentage of students who received Cs? Express your answer to the ones place.

**CHALLENGE QUESTIONS**

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

1.57 Classify each of the following as an observation (O), a hypothesis (H), or an experiment (E): (1.2)
   a. The bicycle tire is flat.
   b. If I add air to the bicycle tire, it will expand to the proper size.
   c. When I added air to the bicycle tire, it was still flat.
   d. The bicycle tire must have a leak in it.

1.58 Classify each of the following as an observation (O), a hypothesis (H), or an experiment (E): (1.2)
   a. A big log in the fire does not burn well.
   b. If I chop the log into smaller wood pieces, it will burn better.
   c. The small wood pieces burn brighter and make a hotter fire.
   d. The small wood pieces are used up faster than burning the big log.

1.59 Solve each of the following for $x$: (1.4)
   a. $10^x = 0.000 001$
   b. $10^x = \frac{1}{1000}$

1.60 Solve each of the following for $z$: (1.4)
   a. $3z + 4 \times 10^{-4} = 0.0124$
   b. $5z - 8 \times 10^{-4} = 6.2 \times 10^{-3}$

1.53 Write each of the following in scientific notation: (1.5)
   a. $1.050 \times 10^5$
   b. $0.000 \times 10^5$
   c. $0.090 \times 10^5$
   d. $2.700 \times 10^5$

1.54 Write each of the following in scientific notation: (1.5)
   a. $0.004 \times 10^3$
   b. $310$
   c. $990 \times 10^3$
   d. $0.000 \times 10^5$

1.55 Write each of the following as a standard number: (1.5)
   a. $2.6 \times 10^{-5}$
   b. $6.5 \times 10^2$
   c. $3.7 \times 10^{-1}$
   d. $5.3 \times 10^5$

1.56 Write each of the following as a standard number: (1.5)
   a. $7.2 \times 10^{-2}$
   b. $1.44 \times 10^3$
   c. $4.8 \times 10^{-4}$
   d. $9.1 \times 10^6$

Use the following graph for problems 1.61 and 1.62:

**ANSWERS**

Answers to Selected Questions and Problems

1.1 a. Chemistry is the study of the composition, structure, properties, and reactions of matter.
   b. A chemical is a substance that has the same composition and properties wherever it is found.

1.3 Many chemicals are listed on a vitamin bottle such as vitamin A, vitamin B₃, vitamin B₁₂, vitamin C, and folic acid.

1.5 Typical items found in a medicine cabinet and some of the chemicals they contain are as follows:
   - Antacid tablets: calcium carbonate, cellulose, starch, stearic acid, silicon dioxide
   - Mouthwash: water, alcohol, thymol, glycerol, sodium benzoate, benzoic acid
   - Cough suppressant: menthol, beta-carotene, sucrose, glucose

1.7 a. A hypothesis proposes a possible explanation for a natural phenomenon.
   b. An experiment is a procedure that tests the validity of a hypothesis.
   c. A theory is a hypothesis that has been validated many times by many scientists.
   d. An observation is a description or measurement of a natural phenomenon.

1.9 a. observation (O)
   b. hypothesis (H)
   c. experiment (E)
   d. observation (O)
   e. observation (O)
   f. conclusion (C)

1.11 a. observation (O)
   b. hypothesis (H)
   c. experiment (E)
   d. experiment (E)
1.13 There are several things you can do that will help you successfully learn chemistry: forming a study group, going to lecture, working Sample Problems and Study Checks, working Questions and Problems and checking answers, reading the assignment ahead of class, and keeping a problem notebook.

1.15 a. c. e

1.17 a. thousandths b. ones c. hundreds

1.19 a. 23 b. –30 c. –2

1.21 a. 84% b. 72% c. 30%

1.23 a. 9 b. 42

1.25 a. The graph shows the relationship between the temperature of a cup of tea and time.
   b. temperature, in °C
   c. 20 °C to 80 °C
   d. decrease

1.27 a. \(5.5 \times 10^4\) b. \(4.8 \times 10^2\) c. \(5 \times 10^{-6}\)
   d. \(1.4 \times 10^{-4}\) e. \(7.2 \times 10^{-3}\) f. \(6.7 \times 10^5\)

1.29 a. 12 000 b. 0.0825
c. 4 000 000 d. 0.0058

1.31 a. \(7.2 \times 10^3\) b. \(3.2 \times 10^{-2}\)
c. \(1 \times 10^4\) d. \(6.8 \times 10^{-2}\)

1.33 27% ethylene glycol

1.35 No. All of the ingredients are chemicals.

1.37 Yes. Sherlock’s investigation includes making observations (gathering data), formulating a hypothesis, testing the hypothesis, and modifying it until one of the hypotheses is validated.

1.39 a. observation (O) b. hypothesis (H)
   c. observation (O) d. conclusion (C)

1.41 a. positive b. positive

1.43 a. observation (O) b. hypothesis (H)
   c. conclusion (C)

1.45 b

1.47 b and e

1.49 a. –2.5 b. –1877 c. 91

1.51 a. 20% b. 40%

1.53 a. \(1.05 \times 10^6\) b. \(3.8 \times 10^{-3}\)
   c. \(9.006 \times 10^{-2}\) d. \(2.7 \times 10^6\)

1.55 a. 0.000 026 b. 650
   c. 0.37 d. 530 000

1.57 a. observation (O) b. hypothesis (H)
   c. experiment (E) d. hypothesis (H)

1.59 a. –6 b. –3

1.61 a. The graph shows the relationship between the solubility of carbon dioxide in water and temperature.
   b. solubility of carbon dioxide (g CO₂/100 g water)
   c. 0 to 0.35 g of CO₂/100 g of water
   d. decrease
Chemistry and Measurements

DURING THE PAST FEW months, Greg has been experiencing an increased number of headaches, and frequently feels dizzy and nauseous. He goes to his doctor’s office where, Sandra, the registered nurse completes the initial part of the exam by recording several measurements: weight 74.8 kg, height 171 cm, temperature 37.2 °C, and blood pressure 155/95. Normal blood pressure is 120/80 or below.

When Greg sees his doctor, he is diagnosed as having high blood pressure (hypertension). The doctor prescribes 80 mg of Inderal (propranolol). Inderal is a beta blocker, which relaxes the muscles of the heart. It is used to treat hypertension, angina (chest pain), arrhythmia, and migraine headaches.

Two weeks later, Greg visits his doctor again, who determines that Greg’s blood pressure is now 152/90. The doctor increases the dosage of Inderal to 160 mg. The registered nurse, Sandra, informs Greg that he needs to increase his daily dosage from two tablets to four tablets.

CAREER
Registered Nurse
In addition to assisting physicians, registered nurses work to promote patient health, and prevent and treat disease. They provide patient care and help patients cope with illness. They take measurements such as a patient’s weight, height, temperature, and blood pressure; make conversions; and calculate drug dosage rates. Registered nurses also maintain detailed medical records of patient symptoms and prescribed medications.
2.1 Units of Measurement

**LEARNING GOAL** Write the names and abbreviations for the metric or SI units used in measurements of length, volume, mass, temperature, and time.

Think about your day. You probably took some measurements. Perhaps you checked your weight by stepping on a bathroom scale. If you made some rice for dinner, you added two cups of water to one cup of rice. If you did not feel well, you may have taken your temperature. Whenever you take a measurement, you use a measuring device such as a scale, a measuring cup, or a thermometer.

Scientists and health professionals throughout the world use the metric system of measurement. It is also the common measuring system in all but a few countries in the world. The International System of Units (SI), or Système International, is the official system of measurement throughout the world except for the United States. In chemistry, we use metric units and SI units for length, volume, mass, temperature, and time, as listed in Table 2.1.

<table>
<thead>
<tr>
<th>Measurement</th>
<th>Metric</th>
<th>SI</th>
</tr>
</thead>
<tbody>
<tr>
<td>Length</td>
<td>meter (m)</td>
<td>meter (m)</td>
</tr>
<tr>
<td>Volume</td>
<td>liter (L)</td>
<td>cubic meter (m³)</td>
</tr>
<tr>
<td>Mass</td>
<td>gram (g)</td>
<td>kilogram (kg)</td>
</tr>
<tr>
<td>Temperature</td>
<td>degree Celsius (°C)</td>
<td>kelvin (K)</td>
</tr>
<tr>
<td>Time</td>
<td>second (s)</td>
<td>second (s)</td>
</tr>
</tbody>
</table>

Suppose you walked 1.3 mi to campus today, carrying a backpack that weighs 26 lb. The temperature was 72 °F. Perhaps you weigh 128 lb and your height is 65 in. These measurements and units may seem familiar to you because they are stated in the U.S. system of measurement. However, in chemistry, we use the metric system in making our measurements. Using the metric system, you walked 2.1 km to campus, carrying a backpack that has a mass of 12 kg, when the temperature was 22 °C. You have a mass of 58.0 kg and a height of 1.7 m.

There are many measurements in everyday life.
2.1 Units of Measurement

Length
The metric and SI unit of length is the **meter (m)**. A meter is 39.37 inches (in.), which makes it slightly longer than a yard (yd). The **centimeter (cm)**, a smaller unit of length, is commonly used in chemistry and is about equal to the width of your little finger. For comparison, there are 2.54 cm in 1 in. (see **FIGURE 2.1**). Some relationships between units for length are

\[
1 \text{ m} = 100 \text{ cm} \\
1 \text{ m} = 39.37 \text{ in.} \\
1 \text{ m} = 1.094 \text{ yd} \\
2.54 \text{ cm} = 1 \text{ in.}
\]

![FIGURE 2.1](image)

Volume
**Volume** \((V)\) is the amount of space a substance occupies. A **liter (L)** is slightly larger than a quart (qt), \((1 \text{ L} = 1.057 \text{ qt})\). The SI unit of volume, the **cubic meter \((m^3)\)** is the volume of a cube that has sides that measure 1 m in length. In a laboratory or a hospital, chemists work with metric units of volume that are smaller and more convenient, such as the **milliliter (mL)**. The volume of 1 mL is the same as 1 cm\(^3\). A liter contains 1000 mL, as shown in **FIGURE 2.2**. A cubic meter is the same volume as 1000 L. Some relationships between units for volume are

\[
1 \text{ m}^3 = 1000 \text{ L} \\
1 \text{ L} = 1000 \text{ mL} \\
1 \text{ mL} = 1 \text{ cm}^3 \\
1 \text{ L} = 1.057 \text{ qt} \\
946.4 \text{ mL} = 1 \text{ qt}
\]

![FIGURE 2.2](image)

Mass
The **mass** of an object is a measure of the quantity of material it contains. The SI unit of mass, the **kilogram (kg)**, is used for larger masses such as body mass. In the metric system, the unit for mass is the **gram (g)**, which is used for smaller masses. There are 1000 g in 1 kg. It takes 2.205 lb to make 1 kg, and 453.6 g is equal to 1 lb. Some relationships between units for mass are

\[
1 \text{ kg} = 1000 \text{ g} \\
1 \text{ kg} = 2.205 \text{ lb} \\
453.6 \text{ g} = 1 \text{ lb}
\]
You may be more familiar with the term weight than with mass. Weight is a measure of the gravitational pull on an object. On the Earth, an astronaut with a mass of 75.0 kg has a weight of 165 lb. On the Moon where the gravitational pull is one-sixth that of the Earth, the astronaut has a weight of 27.5 lb. However, the mass of the astronaut is the same as on Earth, 75.0 kg. Scientists measure mass rather than weight because mass does not depend on gravity.

In a chemistry laboratory, an electronic balance is used to measure the mass in grams of a substance (see FIGURE 2.3).

**Temperature**

Temperature tells us how hot something is, how cold it is outside, or helps us determine if we have a fever (see FIGURE 2.4). In the metric system, temperature is measured using Celsius temperature. On the Celsius (°C) temperature scale, water freezes at 0 °C and boils at 100 °C, whereas on the Fahrenheit (°F) scale, water freezes at 32 °F and boils at 212 °F. In the SI system, temperature is measured using the Kelvin (K) temperature scale on which the lowest possible temperature is 0 K. A unit on the Kelvin scale is called a kelvin (K) and is not written with a degree sign.

**Time**

We typically measure time in units such as years (yr), days, hours (h), minutes (min), or seconds (s). Of these, the SI and metric unit of time is the second (s). The standard now used to determine a second is an atomic clock. Some relationships between units for time are

\[
1 \text{ day} = 24 \text{ h} \\
1 \text{ h} = 60 \text{ min} \\
1 \text{ min} = 60 \text{ s}
\]

**SAMPLE PROBLEM 2.1 Units of Measurement**

On a typical day, a nurse encounters several situations involving measurement. State the name and type of measurement indicated by the units in each of the following:

- a. A patient has a temperature of 38.5 °C.
- b. A physician orders 1.5 g of cefuroxime for injection.
- c. A physician orders 1 L of a sodium chloride solution to be given intravenously.
- d. A medication is to be given to a patient every 4 h.
2.2 Measured Numbers and Significant Figures

LEARNING GOAL Identify a number as measured or exact; determine the number of significant figures in a measured number.

When you make a measurement, you use some type of measuring device. For example, you may use a meterstick to measure your height, a scale to check your weight, or a thermometer to take your temperature.
Measured Numbers

Measured numbers are the numbers you obtain when you measure a quantity such as your height, weight, or temperature. Suppose you are going to measure the lengths of the objects in Figure 2.5. The metric ruler that you use may have lines marked in 1-cm divisions or perhaps in divisions of 0.1 cm. To report the length of the object, you observe the numerical values of the marked lines at the end of the object. Then you can estimate by visually dividing the space between the marked lines. This estimated value is the final digit in a measured number.

For example, in Figure 2.5a, the end of the object is between the marks of 4 cm and 5 cm, which means that the length is more than 4 cm but less than 5 cm. This is written as 4 cm plus an estimated digit. If you estimate that the end of the object is halfway between 4 cm and 5 cm, you would report its length as 4.5 cm. The last digit in a measured number may differ because people do not always estimate in the same way. Thus, someone else might report the length of the same object as 4.4 cm.

The metric ruler shown in Figure 2.5b is marked at every 0.1 cm. With this ruler, you can estimate to the hundredths place (0.01 cm). Now you can determine that the end of the object is between 4.5 cm and 4.6 cm. Perhaps you report its length as 4.55 cm, whereas another student reports its length as 4.56 cm. Both results are acceptable.

In Figure 2.5c, the end of the object appears to line up with the 3-cm mark. Because the divisions are marked in units of 1 cm, the length of the object is between 3 cm and 4 cm. Because the end of the object is on the 3-cm mark, the estimated digit is 0, which means the measurement is reported as 3.0 cm.

Significant Figures

In a measured number, the significant figures (SFs) are all the digits including the estimated digit. Nonzero numbers are always counted as significant figures. However, a zero may or may not be a significant figure depending on its position in a number. Table 2.2 gives the rules and examples of counting significant figures.

<table>
<thead>
<tr>
<th>Rule</th>
<th>Measured Number</th>
<th>Number of Significant Figures</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. A number is a significant figure if it is</td>
<td></td>
<td></td>
</tr>
<tr>
<td>a. not a zero</td>
<td>4.5 g</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>122.35 m</td>
<td>5</td>
</tr>
<tr>
<td>b. a zero between nonzero digits</td>
<td>205 m</td>
<td>3</td>
</tr>
<tr>
<td></td>
<td>5.008 kg</td>
<td>4</td>
</tr>
<tr>
<td>c. a zero at the end of a decimal number</td>
<td>50. L</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>25.0 °C</td>
<td>3</td>
</tr>
<tr>
<td></td>
<td>16.00 g</td>
<td>4</td>
</tr>
<tr>
<td>d. in the coefficient of a number written in scientific notation</td>
<td>$4.8 \times 10^5$ m</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>$5.70 \times 10^{-3}$ g</td>
<td>3</td>
</tr>
<tr>
<td>2. A zero is not significant if it is</td>
<td></td>
<td></td>
</tr>
<tr>
<td>a. at the beginning of a decimal number</td>
<td>0.0004 s</td>
<td>1</td>
</tr>
<tr>
<td></td>
<td>0.075 cm</td>
<td>2</td>
</tr>
<tr>
<td>b. used as a placeholder in a large number without a decimal point</td>
<td>850 000 m</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>1 250 000 g</td>
<td>3</td>
</tr>
</tbody>
</table>

Significant Zeros and Scientific Notation

When one or more zeros in a large number are significant, they are shown clearly by writing the number in scientific notation. For example, if the first zero in the measurement 300 m is significant, but the second zero is not, the measurement is written as $3.0 \times 10^2$ m. In this text, we will place a decimal point after a significant zero at the end of a number. For example, if a measurement is written as 500. g, the decimal point after the second zero
indicates that both zeros are significant. To show this more clearly, we can write it as \(5.00 \times 10^2\) g. We will assume that zeros at the end of large standard numbers without a decimal point are not significant. Therefore, we write 400 000 g as \(4 \times 10^5\) g, which has only one significant figure.

**SAMPLE PROBLEM 2.2 Significant Zeros**

Identify the significant zeros in each of the following measured numbers:

a. 0.000 250 m  
b. 70.040 g  
c. 1 020 000 L

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>measured number</td>
<td>identify significant zeros</td>
<td>rules for significant figures</td>
<td></td>
</tr>
</tbody>
</table>

a. The zero in the last decimal place following the 5 is significant. The zeros preceding the 2 are not significant.  
b. Zeros between nonzero digits or at the end of decimal numbers are significant. All the zeros in 70.040 g are significant.  
c. Zeros between nonzero digits are significant. Zeros at the end of a large number with no decimal point are placeholders but not significant. The zero between 1 and 2 is significant, but the four zeros following the 2 are not significant.

**STUDY CHECK 2.2**

Identify the significant zeros in each of the following measured numbers:

a. 0.040 08 m  
b. \(6.00 \times 10^3\) g

**ANSWER**

a. The zeros between 4 and 8 are significant. The zeros preceding the first nonzero digit 4 are not significant.  
b. All zeros in the coefficient of a number written in scientific notation are significant.

**Exact Numbers**

Exact numbers are those numbers obtained by counting items or using a definition that compares two units in the same measuring system. Suppose a friend asks you how many classes you are taking. You would answer by counting the number of classes in your schedule. It would not use any measuring tool. Suppose you are asked to state the number of seconds in one minute. Without using any measuring device, you would give the definition: There are 60 s in 1 min. Exact numbers are not measured, do not have a limited number of significant figures, and do not affect the number of significant figures in a calculated answer. For more examples of exact numbers, see Table 2.3.

**Table 2.3 Examples of Some Exact Numbers**

<table>
<thead>
<tr>
<th>Counted Numbers</th>
<th>Metric System</th>
<th>U.S. System</th>
</tr>
</thead>
<tbody>
<tr>
<td>8 doughnuts</td>
<td>1 L = 1000 mL</td>
<td>1 ft = 12 in.</td>
</tr>
<tr>
<td>2 baseballs</td>
<td>1 m = 100 cm</td>
<td>1 qt = 4 cups</td>
</tr>
<tr>
<td>5 capsules</td>
<td>1 kg = 1000 g</td>
<td>1 lb = 16 oz</td>
</tr>
</tbody>
</table>

For example, a mass of 42.2 g and a length of \(5.0 \times 10^{-3}\) cm are measured numbers because they are obtained using measuring tools. There are three SFs in 42.2 g because all nonzero digits are always significant. There are two SFs in \(5.0 \times 10^{-3}\) cm because all the
digits in the coefficient of a number written in scientific notation are significant. However, a quantity of three eggs is an exact number that is obtained by counting the eggs. In the equality 1 kg = 1000 g, the masses of 1 kg and 1000 g are both exact numbers because this equality is a definition in the metric system of measurement.

**SAMPLE PROBLEM 2.3 Measured and Exact Numbers**

Identify each of the following numbers as measured or exact and give the number of significant figures (SFs) in each of the measured numbers:

- a. 0.170 L
- b. 4 knives
- c. $6.3 \times 10^{-6}$ s
- d. $1 \text{ m} = 100 \text{ cm}$

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>numbers</td>
<td>identify exact or measured, number of SFs</td>
<td>rules for SFs in measured numbers</td>
<td></td>
</tr>
</tbody>
</table>

a. The volume of 0.170 L is a measured number because it is obtained with a measuring tool. There are three SFs in 0.170 L because nonzero digits are always significant and a zero after nonzero digits in a decimal number is significant.

b. The value of 4 knives is an exact number because it is obtained by counting rather than using a measuring tool.

c. The time of $6.3 \times 10^{-6}$ s is a measured number because it is obtained with a measuring tool. There are two SFs in $6.3 \times 10^{-6}$ s because all the numbers in the coefficient of a number written in scientific notation are significant.

d. The lengths of 1 m and 100 cm are both exact numbers because the relationship $1 \text{ m} = 100 \text{ cm}$ is a definition in the metric system.

**STUDY CHECK 2.3**

Identify each of the following numbers as measured or exact and give the number of significant figures (SFs) in each of the measured numbers:

- a. 0.020 80 kg
- b. $5.06 \times 10^4$ h
- c. 4 chemistry books

**ANSWER**

- a. measured; four SFs
- b. measured; three SFs
- c. exact

**QUICK CHECK 2.3**

Identify each of the following numbers as measured or exact:

- a. A patient has a mass of 67.5 kg.
- b. A patient is given 2 tablets of medication.
- c. In the metric system, 1 L is equal to 1000 mL.
- d. The distance from Denver, Colorado, to Houston, Texas, is 1720 km.

**2.9** What is the estimated digit in each of the following measured numbers?

- a. 8.6 m
- b. 45.25 g
- c. 29 °C

**2.10** What is the estimated digit in each of the following measured numbers?

- a. 125.04 g
- b. 5.057 m
- c. 525.8 °C

**2.11** Identify the numbers in each of the following statements as measured or exact:

- a. A patient has a mass of 67.5 kg.
- b. A patient is given 2 tablets of medication.
- c. In the metric system, 1 L is equal to 1000 mL.
- d. The distance from Denver, Colorado, to Houston, Texas, is 1720 km.

**2.12** Identify the numbers in each of the following statements as measured or exact:

- a. There are 31 students in the laboratory.
- b. The oldest known flower lived $1.20 \times 10^8$ yr ago.
- c. The largest gem ever found, an aquamarine, has a mass of 104 kg.
- d. A laboratory test shows a blood cholesterol level of 184 mg/dL.
2.13 Identify the measured number(s), if any, in each of the following pairs of numbers:
   a. 3 hamburgers and 6 oz of hamburger
   b. 1 table and 4 chairs
   c. 0.75 lb of grapes and 350 g of butter
   d. 60 s = 1 min
2.14 Identify the exact number(s), if any, in each of the following pairs of numbers:
   a. 5 pizzas and 50.0 g of cheese
   b. 6 nickels and 16 g of nickel
   c. 3 onions and 3 lb of onions
   d. 5 miles and 5 cars
2.15 Indicate if the zeros are significant in each of the following measurements:
   a. 0.0038 m
   b. 5.04 cm
   c. 800. L
   d. 3.0 × 10^-3 kg
   e. 85 000 g
2.16 Indicate if the zeros are significant in each of the following measurements:
   a. 20.05 °C
   b. 5.00 m
   c. 0.000 02 g
   d. 120 000 yr
   e. 8.05 × 10^2 L
2.17 How many significant figures are in each of the following?
   a. 11.005 g
   b. 0.000 32 m
   c. 36 000 000 km
   d. 1.80 × 10^4 kg
   e. 0.8250 L
   f. 30.0 °C
2.18 How many significant figures are in each of the following?
   a. 20.60 mL
   b. 1036.48 kg
   c. 4.00 m
   d. 20.8 °C
   e. 60 800 000 g
   f. 5.0 × 10^-3 L

2.19 In which of the following pairs do both numbers contain the same number of significant figures?
   a. 11.0 m and 11.00 m
   b. 0.0250 m and 0.205 m
   c. 0.000 12 s and 12 000 s
   d. 250.0 L and 2.5 × 10^-2 L
2.20 In which of the following pairs do both numbers contain the same number of significant figures?
   a. 0.005 75 g and 5.75 × 10^-3 g
   b. 405 K and 405.0 K
   c. 150 000 s and 1.50 × 10^4 s
   d. 3.8 × 10^-2 L and 3.0 × 10^2 L
2.21 Write each of the following in scientific notation with two significant figures:
   a. 5000 L
   b. 30 000 g
   c. 100 000 m
   d. 0.000 25 cm
2.22 Write each of the following in scientific notation with two significant figures:
   a. 5 100 000 g
   b. 26 000 s
   c. 40 000 m
   d. 0.000 820 kg

Applications
2.23 Identify the number of significant figures in each of the following:
   a. The mass of a neonate is 1.607 kg.
   b. The Daily Value (DV) for iodine for an infant is 130 mcg.
   c. There are 4.02 × 10^4 red blood cells in a blood sample.
2.24 Identify the number of significant figures in each of the following:
   a. An adult with the flu has a temperature of 103.5 °F.
   b. A brain contains 1.20 × 10^10 neurons.
   c. The time for a nerve impulse to travel from the feet to the brain is 0.46 s.

2.3 Significant Figures in Calculations

LEARNING GOAL Adjust calculated answers to give the correct number of significant figures.

In the sciences, we measure many things: the length of a bacterium, the volume of a gas sample, the temperature of a reaction mixture, or the mass of iron in a sample. The numbers obtained from these types of measurements are often used in calculations. The number of significant figures in the measured numbers determines the number of significant figures in the calculated answer.

Using a calculator will help you perform calculations faster. However, calculators cannot think for you. It is up to you to enter the numbers correctly, press the correct function keys, and give an answer with the correct number of significant figures.

Rounding Off

Suppose you decide to buy carpeting for a room that measures 5.52 m by 3.58 m. Each of these measurements has three significant figures because the measuring tape limits your estimated place to 0.01 m. To determine how much carpeting you need, you would calculate the area of the room by multiplying 5.52 times 3.58 on your calculator. The calculator shows the number 19.7616 in its display. However, this is not the correct final answer because there are too many digits, which is the result of the multiplication process. Because each of the original measurements has only three significant figures, the displayed number (19.7616) must be rounded off to three significant figures, 19.8. Therefore, you can order carpeting that will cover an area of 19.8 m².
Each time you use a calculator, it is important to look at the original measurements and determine the number of significant figures that can be used for the answer. You can use the following rules to round off the numbers shown in a calculator display.

### Rules for Rounding Off

1. If the first digit to be dropped is 4 or less, then it and all following digits are simply dropped from the number.
2. If the first digit to be dropped is 5 or greater, then the last retained digit of the number is increased by 1.

<table>
<thead>
<tr>
<th>Number to Round Off</th>
<th>Three Significant Figures</th>
<th>Two Significant Figures</th>
</tr>
</thead>
<tbody>
<tr>
<td>8.4234</td>
<td>8.42 (drop 34)</td>
<td>8.4 (drop 234)</td>
</tr>
<tr>
<td>14.780</td>
<td>14.7 (drop 80, increase the last retained digit by 1)</td>
<td>15 (drop 780, increase the last retained digit by 1)</td>
</tr>
<tr>
<td>3256</td>
<td>3260 (drop 6, increase the last retained digit by 1, add 0) (3.26 × 10³)</td>
<td>3300 (drop 56, increase the last retained digit by 1, add 00) (3.3 × 10³)</td>
</tr>
</tbody>
</table>

*The value of a large number is retained by using placeholder zeros to replace dropped digits.

### SAMPLE PROBLEM 2.4 Rounding Off

Round off each of the following numbers to three significant figures:

a. 35.7823 m  
   b. 0.002 621 7 L  
   c. 3.8268 × 10³ g

#### TRY IT FIRST

**SOLUTION**

a. To round off 35.7823 m to three significant figures, drop 823 and increase the last retained digit by 1 to give 35.8 m.

b. To round off 0.002 621 7 L to three significant figures, drop 17 to give 0.002 62 L.

c. To round off 3.8268 × 10³ g to three significant figures, drop 68 and increase the last retained digit by 1 to give 3.83 × 10³ g.

#### STUDY CHECK 2.4

Round off each of the numbers in Sample Problem 2.4 to two significant figures.

**ANSWER**

a. 36 m  
   b. 0.0026 L  
   c. 3.8 × 10³ g

### Multiplication and Division with Measured Numbers

In multiplication or division, the final answer is written so that it has the same number of significant figures (SFs) as the measurement with the fewest SFs. An example of rounding off a calculator display follows:

Perform the following operations with measured numbers:

\[
\frac{2.8 \times 67.40}{34.8} = \text{5.422980506}
\]

When the problem has multiple steps, the numbers in the numerator are multiplied and then divided by each of the numbers in the denominator.

\[
\begin{align*}
2.8 & \times 67.40 \div 34.8 & = & \text{5.422980506} \\
\text{Two SFs} & \text{ Four SFs} & \text{ Three SFs} & \text{ Calculator display} & \text{ Answer, rounded off to two SFs}
\end{align*}
\]
Because the calculator display has more digits than the significant figures in the measured numbers allow, we need to round off. Using the measured number that has the fewest number (two) of significant figures, 2.8, we round off the calculator display to the answer with two SFs.

**Adding Significant Zeros**

Sometimes, a calculator display gives a small whole number. Then we add one or more significant zeros to the calculator display to obtain the correct number of significant figures. For example, suppose the calculator display is 4, but you used measurements that have three significant numbers. Then two significant zeros are added to give 4.00 as the correct answer.

\[
\begin{align*}
\text{Three SFs} & \quad \frac{8.00}{2.00} = 4 \quad \text{Final answer, two zeros added to give three SFs} \\
\end{align*}
\]

**SAMPLE PROBLEM 2.5 Significant Figures in Multiplication and Division**

Perform the following calculations with measured numbers. Round off the calculator display or add zeros to give each answer with the correct number of significant figures.

a. \(56.8 \times 0.37\)  

b. \((2.075)(0.585)\)  

c. \(25.0 \div (5.00)\)

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>multiplication and division</td>
<td>answer with SFs</td>
<td>rules for rounding off, adding zeros</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Determine the number of significant figures in each measured number.

a. Three SFs  

b. Four SFs  

c. Three SFs

**STEP 2** Perform the indicated calculation.

a. \(21.016\)  

b. \(5.884313345\)  

c. \(5.00\)

**STEP 3** Round off (or add zeros) to give the same number of significant figures as the measurement having the fewest significant figures.

a. 21 (rounded off to two SFs)  

b. 5.88 (rounded off to three SFs)  

c. 5.00 (two zeros added)

**STUDY CHECK 2.5**

Perform the following calculations with measured numbers and give the answers with the correct number of significant figures:

a. \(45.26 \times 0.010\)  

b. \(2.6 \div 324\)  

c. \(4.0 \times 8.0\)  

**ANSWER**

a. 0.4924  

b. 0.0080 or \(8.0 \times 10^{-3}\)  

c. 2.0
Addition and Subtraction with Measured Numbers

In addition or subtraction, the final answer is written so that it has the same number of decimal places as the measurement having the fewest decimal places.

2.045 Thousandths place
+ 34.1 Tenths place
\[ \boxed{36.145} \] Calculator display
36.1 Answer, rounded off to the tenths place

When numbers are added or subtracted to give an answer ending in zero, the zero does not appear after the decimal point in the calculator display. For example, \(14.5 \text{ g} - 2.5 \text{ g} = 12.0 \text{ g}\). However, if you do the subtraction on your calculator, the display shows 12. To write the correct answer, a significant zero is written after the decimal point.

SAMPLE PROBLEM 2.6 Significant Figures in Addition and Subtraction

Perform the following calculations and give each answer with the correct number of decimal places:

a. \(104.45 \text{ mL} + 0.838 \text{ mL} + 46 \text{ mL}\)

b. \(153.247 \text{ g} - 14.82 \text{ g}\)

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>addition</td>
<td>correct number</td>
<td>rules for rounding off</td>
</tr>
<tr>
<td>subtraction</td>
<td>of decimal places</td>
<td></td>
</tr>
</tbody>
</table>

STEP 1 Determine the number of decimal places in each measured number.

a. 104.45 mL Hundredths place
b. 153.247 g Thousandths place

0.838 mL Thousandths place

46 mL Ones place

STEP 2 Perform the indicated calculation.

a. \(151.288\) Calculator display

b. \(138.427\) Calculator display

STEP 3 Round off the answer to give the same number of decimal places as the measurement having the fewest decimal places.

a. 151 mL (rounded off to the ones place)

b. 138.43 g (rounded off to the hundredths place)

STUDY CHECK 2.6

Perform the following calculations and give each answer with the correct number of decimal places:

a. \(82.45 \text{ mg} + 1.245 \text{ mg} + 0.000 56 \text{ mg}\)

b. \(4.259 \text{ L} - 3.8 \text{ L}\)

ANSWER

a. 83.70 mg

b. 0.5 L
2.4 Prefixes and Equalities

LEARNING GOAL Use the numerical values of prefixes to write a metric equality.

The special feature of the SI as well as the metric system is that a prefix can be placed in front of any unit to increase or decrease its size by some factor of 10. For example, the prefixes milli and micro are used to make the smaller units, milligram (mg) and microgram (μg).

The U.S. Food and Drug Administration has determined the Daily Values (DV) for nutrients for adults and children aged 4 or older. Examples of these recommended Daily Values, some of which use prefixes, are listed in Table 2.4.

The prefix centi is like cents in a dollar. One cent would be a “centidollar” or 0.01 of a dollar. That also means that one dollar is the same as 100 cents. The prefix deci is like dimes in a dollar. One dime would be a “decidollar” or 0.1 of a dollar. That also means that one dollar is the same as 10 dimes.

The relationship of a prefix to a unit can be expressed by replacing the prefix with its numerical value. For example, when the prefix kilo in kilometer is replaced with its value of 1000, we find that a kilometer is equal to 1000 m. Other examples follow:

1 kilometer (1 km) = 1000 meters (1000 m = 10^3 m)
1 kiloliter (1 kL) = 1000 liters (1000 L = 10^3 L)
1 kilogram (1 kg) = 1000 grams (1000 g = 10^3 g)

### Table 2.4: Daily Values for Selected Nutrients

<table>
<thead>
<tr>
<th>Nutrient</th>
<th>Amount Recommended</th>
</tr>
</thead>
<tbody>
<tr>
<td>Protein</td>
<td>50 g</td>
</tr>
<tr>
<td>Vitamin C</td>
<td>60 mg</td>
</tr>
<tr>
<td>Vitamin B12</td>
<td>6 μg (6 mcg)</td>
</tr>
<tr>
<td>Calcium</td>
<td>1000 mg</td>
</tr>
<tr>
<td>Copper</td>
<td>2 mg</td>
</tr>
<tr>
<td>Iodine</td>
<td>150 μg (150 mcg)</td>
</tr>
<tr>
<td>Iron</td>
<td>18 mg</td>
</tr>
<tr>
<td>Magnesium</td>
<td>400 mg</td>
</tr>
<tr>
<td>Niacin</td>
<td>20 mg</td>
</tr>
<tr>
<td>Potassium</td>
<td>4700 mg</td>
</tr>
<tr>
<td>Selenium</td>
<td>70 μg (70 mcg)</td>
</tr>
<tr>
<td>Sodium</td>
<td>2400 mg</td>
</tr>
<tr>
<td>Zinc</td>
<td>15 mg</td>
</tr>
</tbody>
</table>
**TABLE 2.5 Metric and SI Prefixes**

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Numerical Value</th>
<th>Scientific Notation</th>
<th>Equality</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Prefixes That Increase the Size of the Unit</strong></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>peta</td>
<td>P</td>
<td>1 000 000 000 000</td>
<td>$10^{15}$</td>
<td>$1 \text{Pg} = 1 \times 10^{15} \text{g}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{g} = 1 \times 10^{-15} \text{Pg}$</td>
</tr>
<tr>
<td>tera</td>
<td>T</td>
<td>1 000 000 000 000</td>
<td>$10^{12}$</td>
<td>$1 \text{Ts} = 1 \times 10^{12} \text{s}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{s} = 1 \times 10^{-12} \text{Ts}$</td>
</tr>
<tr>
<td>giga</td>
<td>G</td>
<td>1 000 000 000</td>
<td>$10^9$</td>
<td>$1 \text{Gm} = 1 \times 10^9 \text{m}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{m} = 1 \times 10^{-9} \text{Gm}$</td>
</tr>
<tr>
<td>mega</td>
<td>M</td>
<td>1 000 000</td>
<td>$10^6$</td>
<td>$1 \text{Mp} = 1 \times 10^6 \text{g}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{g} = 1 \times 10^{-6} \text{Mp}$</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>1 000</td>
<td>$10^3$</td>
<td>$1 \text{km} = 1 \times 10^3 \text{m}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{m} = 1 \times 10^{-3} \text{km}$</td>
</tr>
<tr>
<td><strong>Prefixes That Decrease the Size of the Unit</strong></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>deci</td>
<td>d</td>
<td>0.1</td>
<td>$10^{-1}$</td>
<td>$1 \text{dL} = 1 \times 10^{-1} \text{L}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{L} = 10 \text{dL}$</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>0.01</td>
<td>$10^{-2}$</td>
<td>$1 \text{cm} = 1 \times 10^{-2} \text{m}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{m} = 100 \text{cm}$</td>
</tr>
<tr>
<td>milli</td>
<td>m</td>
<td>0.001</td>
<td>$10^{-3}$</td>
<td>$1 \text{ms} = 1 \times 10^{-3} \text{s}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{s} = 1 \times 10^3 \text{ms}$</td>
</tr>
<tr>
<td>micro</td>
<td>µ</td>
<td>0.000001</td>
<td>$10^{-6}$</td>
<td>$1 \mu\text{g} = 1 \times 10^{-6} \mu\text{g}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \mu\text{g} = 1 \times 10^{6} \mu\text{g}$</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>0.00000001</td>
<td>$10^{-9}$</td>
<td>$1 \text{nm} = 1 \times 10^{-9} \text{m}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{m} = 1 \times 10^9 \text{nm}$</td>
</tr>
<tr>
<td>pico</td>
<td>p</td>
<td>0.0000000001</td>
<td>$10^{-12}$</td>
<td>$1 \text{ps} = 1 \times 10^{-12} \text{s}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{s} = 1 \times 10^{12} \text{ps}$</td>
</tr>
<tr>
<td>femto</td>
<td>f</td>
<td>0.00000000001</td>
<td>$10^{-15}$</td>
<td>$1 \text{fs} = 1 \times 10^{-15} \text{s}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$1 \text{s} = 1 \times 10^{15} \text{fs}$</td>
</tr>
</tbody>
</table>

*In medicine, the abbreviation mc for the prefix micro is used because the symbol µ may be misread, which could result in a medication error. Thus, 1 µg would be written as 1 mcg.*

---

**SAMPLE PROBLEM 2.7 Prefixes and Equalities**

An endoscopic camera has a width of 1 mm. Complete each of the following equalities involving millimeters:

**a.** $1 \text{m} = \underline{\phantom{1}} \text{mm}$  
**b.** $1 \text{cm} = \underline{\phantom{1}} \text{mm}$

**TRY IT FIRST**

**SOLUTION**

**a.** $1 \text{m} = 1000 \text{mm}$  
**b.** $1 \text{cm} = 10 \text{mm}$

**STUDY CHECK 2.7**

What is the relationship between millimeters and micrometers?

**ANSWER**

$1 \text{mm} = 1000 \mu\text{m} \text{ (mcm)}$

---

**Measuring Length**

An ophthalmologist may measure the diameter of the retina of an eye in centimeters (cm), whereas a surgeon may need to know the length of a nerve in millimeters (mm). When the
prefix *centi* is used with the unit meter, it becomes *centimeter*, a length that is one-hundredth of a meter (0.01 m). When the prefix *milli* is used with the unit meter, it becomes *millimeter*, a length that is one-thousandth of a meter (0.001 m). There are 100 cm and 1000 mm in a meter.

If we compare the lengths of a millimeter and a centimeter, we find that 1 mm is 0.1 cm; there are 10 mm in 1 cm. These comparisons are examples of *equalities*, which show the relationship between two units that measure the same quantity. In every equality, each quantity has both a number and a unit. Examples of equalities between different metric units of length follow:

\[
\begin{align*}
1 \text{ m} &= 100 \text{ cm} = 1 \times 10^2 \text{ cm} \\
1 \text{ m} &= 1000 \text{ mm} = 1 \times 10^3 \text{ mm} \\
1 \text{ cm} &= 10 \text{ mm} = 1 \times 10^1 \text{ mm}
\end{align*}
\]

Some metric units for length are compared in FIGURE 2.6.

![FIGURE 2.6](image)

Using a retinal camera, an ophthalmologist photographs the retina of an eye.

This example of an equality shows the relationship between meters and centimeters.

Equality: \[1 \text{ m} = 100 \text{ cm}\]

**Measuring Volume**

Volumes of 1 L or smaller are common in the health sciences. When a liter is divided into 10 equal portions, each portion is a deciliter (dL). There are 10 dL in 1 L. Laboratory results for blood work are often reported in mass per deciliter. **TABLE 2.6** lists typical laboratory test values for some substances in the blood.

**TABLE 2.6 Some Typical Laboratory Test Values**

<table>
<thead>
<tr>
<th>Substance in Blood</th>
<th>Typical Range</th>
</tr>
</thead>
<tbody>
<tr>
<td>Albumin</td>
<td>3.5–5.0 g/dL</td>
</tr>
<tr>
<td>Ammonia</td>
<td>20–70 µg/dL (mcg/dL)</td>
</tr>
<tr>
<td>Calcium</td>
<td>8.5–10.5 mg/dL</td>
</tr>
<tr>
<td>Cholesterol</td>
<td>105–250 mg/dL</td>
</tr>
<tr>
<td>Iron (male)</td>
<td>80–160 µg/dL (mcg/dL)</td>
</tr>
<tr>
<td>Protein (total)</td>
<td>6.0–8.0 g/dL</td>
</tr>
</tbody>
</table>
When a liter is divided into a thousand parts, each of the smaller volumes is called a milliliter (mL). In a 1-L container of physiological saline, there are 1000 mL of solution. In 1 mL of a laboratory sample, there are 1000 mcL (see FIGURE 2.7). Examples of equalities between different metric units of volume follow:

\[
\begin{align*}
1 \text{ L} &= 10 \text{ dL} = 1 \times 10^1 \text{ dL} \\
1 \text{ L} &= 1000 \text{ mL} = 1 \times 10^3 \text{ mL} \\
1 \text{ dL} &= 100 \text{ mL} = 1 \times 10^2 \text{ mL} \\
1 \text{ mL} &= 1000 \mu\text{L (mcL)} = 1 \times 10^3 \mu\text{L (mcL)}
\end{align*}
\]

The cubic centimeter (abbreviated as cm\(^3\) or cc) is the volume of a cube whose dimensions are 1 cm on each side. A cubic centimeter has the same volume as a milliliter, and the units are often used interchangeably.

\[
1 \text{ cm}^3 = 1 \text{ cc} = 1 \text{ mL}
\]

When you see 1 cm, you are reading about length; when you see 1 cm\(^3\) or 1 cc or 1 mL, you are reading about volume. A comparison of units of volume is illustrated in FIGURE 2.8.

**Measuring Mass**

When you go to the doctor for a physical examination, your mass is recorded in kilograms, whereas the results of your laboratory tests are reported in grams, milligrams (mg), or micrograms (\(\mu\text{g}\) or mcg). A kilogram is equal to 1000 g. One gram represents the same mass as 1000 mg, and one mg equals 1000 \(\mu\text{g}\) (or 1000 mcg). Examples of equalities between different metric units of mass follow:

\[
\begin{align*}
1 \text{ kg} &= 1000 \text{ g} = 1 \times 10^3 \text{ g} \\
1 \text{ g} &= 1000 \text{ mg} = 1 \times 10^3 \text{ mg} \\
1 \text{ mg} &= 1000 \mu\text{g (mcg)} = 1 \times 10^3 \mu\text{g (mcg)}
\end{align*}
\]
2.39 Write the complete name for each of the following units:
   a. cL  b. kg  c. ms  d. Gm

2.40 Write the complete name for each of the following units:
   a. dL  b. Ts  c. mcg  d. pm

2.41 Write the numerical value for each of the following prefixes:
   a. centi  b. tera  c. milli  d. deci

2.42 Write the numerical value for each of the following prefixes:
   a. giga  b. micro  c. mega  d. nano

2.43 Use a prefix to write the name for each of the following:
   a. 0.1 g  b. 10⁻⁶ g  c. 1000 g  d. 0.01 g

2.44 Use a prefix to write the name for each of the following:
   a. 10³ m  b. 10⁶ m  c. 0.001 m  d. 10⁻¹⁵ m

2.45 Complete each of the following metric relationships:
   a. 1 m = ________ cm  b. 1 m = ________ mm
   c. 1 mm = ________ m  d. 1 L = ________ mL

2.46 Complete each of the following metric relationships:
   a. 1 Mg = ________ g  b. 1 mL = ________ µL
   c. 1 g = ________ kg  d. 1 g = ________ mg

2.47 For each of the following pairs, which is the larger unit?
   a. milligram or kilogram  b. milliliter or microliter
   c. m or km  d. kL or dL
   e. nanometer or picometer

2.48 For each of the following pairs, which is the smaller unit?
   a. mg or g  b. centimeter or nanometer
   c. millimeter or micrometer  d. mL or dL

2.5 Writing Conversion Factors

LEARNING GOAL Write a conversion factor for two units that describe the same quantity.

Many problems in chemistry and the health sciences require you to change from one unit to another unit. You make changes in units every day. For example, suppose you worked 2.0 h on your homework, and someone asked you how many minutes that was. You would answer 120 min. You must have multiplied 2.0 h × 60 min/h because you knew the equality (1 h = 60 min) that related the two units. When you expressed 2.0 h as 120 min, you did not change the amount of time you spent studying. You changed only the unit of measurement used to express the time. Any equality can be written as fractions called conversion factors with one of the quantities in the numerator and the other quantity in the denominator. Be sure to include the units when you write the conversion factors. Two conversion factors are always possible from any equality.

**Two Conversion Factors for the Equality: 1 h = 60 min**

\[
\frac{\text{Numerator}}{\text{Denominator}} \rightarrow \frac{60 \text{ min}}{1 \text{ h}} \quad \text{and} \quad \frac{1 \text{ h}}{60 \text{ min}}
\]

These factors are read as “60 minutes per 1 hour” and “1 hour per 60 minutes.” The term per means “divide.” Some common relationships are given in Table 2.7. It is important that the equality you select to form a conversion factor is a true relationship.

The numbers in any equality between two metric units or between two U.S. system units are obtained by definition. Because numbers in a definition are exact, they are not used to determine significant figures. For example, the equality of 1 g = 1000 mg is
defined, which means that both of the numbers 1 and 1000 are exact. However, when an equality consists of a metric unit and a U.S. unit, one of the numbers in the equality is obtained by measurement and counts toward the significant figures in the answer. For example, the equality of 1 lb = 453.6 g is obtained by measuring the grams in exactly 1 lb. In this equality, the measured quantity 453.6 g has four significant figures, whereas the 1 is exact. An exception is the relationship of 1 in. = 2.54 cm, which has been defined as exact.

Metric Conversion Factors

We can write metric conversion factors for any of the metric relationships. For example, from the equality for meters and centimeters, we can write the following factors:

<table>
<thead>
<tr>
<th>Metric Equality</th>
<th>Conversion Factors</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 m = 100 cm</td>
<td>$\frac{100 \text{ cm}}{1 \text{ m}}$ and $\frac{1 \text{ m}}{100 \text{ cm}}$</td>
</tr>
</tbody>
</table>

Both are proper conversion factors for the relationship; one is just the inverse of the other. The usefulness of conversion factors is enhanced by the fact that we can turn a conversion factor over and use its inverse. The numbers 100 and 1 in this equality and its conversion factors are both exact numbers.

Metric–U.S. System Conversion Factors

Suppose you need to convert from pounds, a unit in the U.S. system, to kilograms in the metric (or SI) system. A relationship you could use is

$$1 \text{ kg} = 2.205 \text{ lb}$$

The corresponding conversion factors would be

$$\frac{2.205 \text{ lb}}{1 \text{ kg}} \text{ and } \frac{1 \text{ kg}}{2.205 \text{ lb}}$$

$\text{FIGURE 2.9}$ illustrates the contents of some packaged foods in both U.S. and metric units.

---

**TABLE 2.7** Some Common Equalities

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Metric (SI)</th>
<th>U.S.</th>
<th>Metric–U.S.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Length</td>
<td>1 km = 1000 m</td>
<td>1 ft = 12 in.</td>
<td>2.54 cm = 1 in. (exact)</td>
</tr>
<tr>
<td></td>
<td>1 m = 1000 mm</td>
<td>1 yd = 3 ft</td>
<td>1 m = 39.37 in.</td>
</tr>
<tr>
<td></td>
<td>1 cm = 10 mm</td>
<td>1 mi = 5280 ft</td>
<td>1 km = 0.6214 mi</td>
</tr>
<tr>
<td>Volume</td>
<td>1 L = 1000 mL</td>
<td>1 qt = 4 cups</td>
<td>946.4 mL = 1 qt</td>
</tr>
<tr>
<td></td>
<td>1 dL = 100 mL</td>
<td>1 qt = 2 pt</td>
<td>1 L = 1.057 qt</td>
</tr>
<tr>
<td></td>
<td>1 mL = 1 cm³</td>
<td>1 gal = 4 qt</td>
<td>473.2 mL = 1 pt</td>
</tr>
<tr>
<td></td>
<td>1 mL = 1 cc⁺</td>
<td>3.785 L = 1 gal</td>
<td>5 mL = 1 tsp⁺</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>15 mL = 1 T (tbsp)⁺</td>
</tr>
<tr>
<td>Mass</td>
<td>1 kg = 1000 g</td>
<td>1 lb = 16 oz</td>
<td>1 kg = 2.205 lb</td>
</tr>
<tr>
<td></td>
<td>1 g = 1000 mg</td>
<td></td>
<td>453.6 g = 1 lb</td>
</tr>
<tr>
<td></td>
<td>1 mg = 1000 mcg⁺</td>
<td></td>
<td>28.35 g = 1 oz</td>
</tr>
<tr>
<td>Time</td>
<td>1 h = 60 min</td>
<td>1 h = 60 min</td>
<td></td>
</tr>
<tr>
<td></td>
<td>1 min = 60 s</td>
<td>1 min = 60 s</td>
<td></td>
</tr>
</tbody>
</table>

*Used in medicine.

---

**ENGAGE**

Why does the equality 1 day = 24 h have two conversion factors?

---

**FIGURE 2.9** In the United States, the contents of many packaged foods are listed in both U.S. and metric units.

What are some advantages of using the metric system?
SAMPLE PROBLEM 2.8 Writing Conversion Factors

Identify the correct conversion factor(s) for the equality: 1 pt of blood contains 473.2 mL of blood.

a. \( \frac{473.2 \text{ pt}}{1 \text{ mL}} \)  

b. \( \frac{1 \text{ pt}}{473.2 \text{ mL}} \)  

c. \( \frac{473.2 \text{ mL}}{1 \text{ pt}} \)  

d. \( \frac{1 \text{ mL}}{473.2 \text{ pt}} \)

TRY IT FIRST

SOLUTION

The equality for pints and milliliters is \( 1 \text{ pt} = 473.2 \text{ mL} \). Answers b and e are correctly written conversion factors for the equality.

STUDY CHECK 2.8

What are the two correctly written conversion factors for the equality: 1000 mm = 1 m?

ANSWER

\( \frac{1000 \text{ mm}}{1 \text{ m}} \) and \( \frac{1 \text{ m}}{1000 \text{ mm}} \)

Conversion Factors with Powers

Sometimes we use a conversion factor that is squared or cubed. This is the case when we need to calculate an area or a volume.

Distance = length
Area = length \times length = length^2
Volume = length \times length \times length = length^3

Suppose you want to write the equality and the conversion factors for the relationship between an area in square centimeters and in square meters. To square the equality \( 1 \text{ m} = 100 \text{ cm} \), we square both the number and the unit on each side.

Equality: \( 1 \text{ m} = 100 \text{ cm} \)
Area: \( (1 \text{ m})^2 = (100 \text{ cm})^2 \) or \( 1 \text{ m}^2 = (100 \text{ cm})^2 \)

From the new equality, we can write two conversion factors as follows:

Conversion factors: \( \frac{(100 \text{ cm})^2}{(1 \text{ m})^2} \) and \( \frac{(1 \text{ m})^2}{(100 \text{ cm})^2} \)

In the following example, we show that the equality \( 1 \text{ in.} = 2.54 \text{ cm} \) can be squared to give area or can be cubed to give volume. Both the number and the unit must be squared or cubed.

<table>
<thead>
<tr>
<th>Measurement</th>
<th>Equality</th>
<th>Conversion Factors</th>
</tr>
</thead>
<tbody>
<tr>
<td>Length</td>
<td>1 in. = 2.54 cm</td>
<td>( \frac{2.54 \text{ cm}}{1 \text{ in.}} ) and ( \frac{1 \text{ in.}}{2.54 \text{ cm}} )</td>
</tr>
</tbody>
</table>
| Area | \( (1 \text{ in.})^2 = (2.54 \text{ cm})^2 \)  
\( (1 \text{ in.})^2 = (2.54)^2 \text{ cm}^2 = 6.45 \text{ cm}^2 \) | \( \frac{(2.54 \text{ cm})^2}{(1 \text{ in.})^2} \) and \( \frac{(1 \text{ in.})^2}{(2.54 \text{ cm})^2} \) |
| Volume | \( (1 \text{ in.})^3 = (2.54 \text{ cm})^3 \)  
\( (1 \text{ in.})^3 = (2.54)^3 \text{ cm}^3 = 16.4 \text{ cm}^3 \) | \( \frac{(2.54 \text{ cm})^3}{(1 \text{ in.})^3} \) and \( \frac{(1 \text{ in.})^3}{(2.54 \text{ cm})^3} \)

Equalities and Conversion Factors Stated Within a Problem

An equality may also be stated within a problem that applies only to that problem. For example, the speed of a car in kilometers per hour or the milligrams of vitamin C in a tablet would be specific relationships for that problem only. However, it is possible to identify these relationships within a problem and to write corresponding conversion factors.
From each of the following statements, we can write an equality, and two conversion factors, and identify each number as exact or give the number of significant figures.

The car was traveling at a speed of 85 km/h.

<table>
<thead>
<tr>
<th>Equality</th>
<th>Conversion Factors</th>
<th>Significant Figures or Exact</th>
</tr>
</thead>
<tbody>
<tr>
<td>$85 \text{ km} = 1 \text{ h}$</td>
<td>$\frac{85 \text{ km}}{1 \text{ h}}$ and $\frac{1 \text{ h}}{85 \text{ km}}$</td>
<td>The 85 km is measured: It has two significant figures. The 1 h is exact.</td>
</tr>
</tbody>
</table>

One tablet contains 500 mg of vitamin C.

<table>
<thead>
<tr>
<th>Equality</th>
<th>Conversion Factors</th>
<th>Significant Figures or Exact</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 tablet $= 500 \text{ mg}$ of vitamin C</td>
<td>$\frac{500 \text{ mg vitamin C}}{1 \text{ tablet}}$ and $\frac{1 \text{ tablet}}{500 \text{ mg vitamin C}}$</td>
<td>The 500 mg is measured: It has one significant figure. The 1 tablet is exact.</td>
</tr>
</tbody>
</table>

Conversion Factors from a Percentage, ppm, and ppb

A percentage ($\%$) is written as a conversion factor by choosing a unit and expressing the numerical relationship of the parts of this unit to 100 parts of the whole. For example, a person might have 18% body fat by mass. The percentage quantity can be written as 18 mass units of body fat in every 100 mass units of body mass. Different mass units such as grams (g), kilograms (kg), or pounds (lb) can be used, but both units in the factor must be the same.

<table>
<thead>
<tr>
<th>Equality</th>
<th>Conversion Factors</th>
<th>Significant Figures or Exact</th>
</tr>
</thead>
<tbody>
<tr>
<td>$18 \text{ kg of body fat} = 100 \text{ kg of body mass}$</td>
<td>$\frac{18 \text{ kg body fat}}{100 \text{ kg body mass}}$ and $\frac{100 \text{ kg body mass}}{18 \text{ kg body fat}}$</td>
<td>The 18 kg is measured: It has two significant figures. The 100 kg is exact.</td>
</tr>
</tbody>
</table>

When scientists want to indicate very small ratios, they use numerical relationships called parts per million (ppm) or parts per billion (ppb). The ratio of parts per million is the same as the milligrams of a substance per kilogram (mg/kg). The ratio of parts per billion equals the micrograms per kilogram ($\mu$g/kg, mcg/kg).

<table>
<thead>
<tr>
<th>Ratio</th>
<th>Units</th>
</tr>
</thead>
<tbody>
<tr>
<td>parts per million (ppm)</td>
<td>milligrams per kilogram (mg/kg)</td>
</tr>
<tr>
<td>parts per billion (ppb)</td>
<td>micrograms per kilogram ($\mu$g/kg, mcg/kg)</td>
</tr>
</tbody>
</table>

For example, the maximum amount of lead that is allowed by the Food and Drug Administration (FDA) in glazed pottery bowls is 2 ppm, which is 2 mg/kg.

<table>
<thead>
<tr>
<th>Equality</th>
<th>Conversion Factors</th>
<th>Significant Figures or Exact</th>
</tr>
</thead>
<tbody>
<tr>
<td>$2 \text{ mg of lead} = 1 \text{ kg of glaze}$</td>
<td>$\frac{2 \text{ mg lead}}{1 \text{ kg glaze}}$ and $\frac{1 \text{ kg glaze}}{2 \text{ mg lead}}$</td>
<td>The 2 mg is measured: It has one significant figure. The 1 kg is exact.</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 2.9** Equalities and Conversion Factors Stated in a Problem

Write the equality and two conversion factors, and identify each number as exact or give the number of significant figures for each of the following:

a. The medication that Greg takes for his high blood pressure contains 40. mg of propranolol in 1 tablet.

b. Cold water fish such as salmon contains 1.9% omega-3 fatty acids by mass.
c. The U.S. Environmental Protection Agency (EPA) has set the maximum level for mercury in tuna at 0.5 ppm.

**TRY IT FIRST**

**SOLUTION**

a. There are 40. mg of propranolol in 1 tablet.

<table>
<thead>
<tr>
<th>Equality</th>
<th>Conversion Factors</th>
<th>Significant Figures or Exact</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 tablet = 40. mg of propranolol</td>
<td>40. mg propranolol</td>
<td>The 40. mg is measured: It has two significant figures. The 1 tablet is exact.</td>
</tr>
</tbody>
</table>

b. Cold water fish such as salmon contains 1.9% omega-3 fatty acids by mass.

<table>
<thead>
<tr>
<th>Equality</th>
<th>Conversion Factors</th>
<th>Significant Figures or Exact</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.9 g of omega-3 fatty acids = 100 g of salmon</td>
<td>1.9 g omega-3 fatty acids = 100 g salmon and 1.9 g omega-3 fatty acids</td>
<td>The 1.9 g is measured: It has two significant figures. The 100 g is exact.</td>
</tr>
</tbody>
</table>

c. The EPA has set the maximum level for mercury in tuna at 0.5 ppm.

<table>
<thead>
<tr>
<th>Equality</th>
<th>Conversion Factors</th>
<th>Significant Figures or Exact</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.5 mg of mercury = 1 kg of tuna</td>
<td>0.5 mg mercury = 1 kg tuna and 0.5 mg mercury</td>
<td>The 0.5 mg is measured: It has one significant figure. The 1 kg is exact.</td>
</tr>
</tbody>
</table>

**STUDY CHECK 2.9**

Levsin (hyoscyamine), used to treat stomach and bladder problems, is available as drops with 0.125 mg Levsin per 1 mL of solution. Write the equality and two conversion factors, and identify each number as exact or give the number of significant figures.

**ANSWER**

0.125 mg of Levsin = 1 mL of solution

0.125 mg Levsin and 1 mL solution = 0.125 mg Levsin

The 0.125 mg is measured: It has three SFs. The 1 mL is exact.

**QUESTIONS AND PROBLEMS**

**2.5 Writing Conversion Factors**

**LEARNING GOAL** Write a conversion factor for two units that describe the same quantity.

2.49 Why can two conversion factors be written for an equality such as 1 m = 100 cm?

2.50 How can you check that you have written the correct conversion factors for an equality?

2.51 Write the equality and two conversion factors for each of the following pairs of units:

a. centimeters and meters
b. nanograms and grams
c. liters and kiloliters
d. seconds and milliseconds
e. cubic meters and cubic centimeters
2.52 Write the equality and two conversion factors for each of the following pairs of units:
   a. centimeters and inches
   b. kilometers and miles
   c. pounds and grams
   d. gallons and liters
   e. square centimeters and square inches

2.53 Write the equality and two conversion factors, and identify the numbers as exact or give the number of significant figures for each of the following:
   a. One yard is 3 ft.
   b. One kilogram is 2.205 lb.
   c. One minute is 60 s.
   d. A car goes 27 mi on 1 gal of gas.
   e. Sterling silver is 93% silver by mass.

2.54 Write the equality and two conversion factors, and identify the numbers as exact or give the number of significant figures for each of the following:
   a. One liter is 1.057 qt.
   b. At the store, oranges are $1.29 per lb.
   c. There are 7 days in 1 week.
   d. One deciliter contains 100 mL.
   e. An 18-carat gold ring contains 75% gold by mass.

2.55 Write the equality and two conversion factors, and identify the numbers as exact or give the number of significant figures for each of the following:
   a. A bee flies at an average speed of 3.5 m per second.
   b. The Daily Value (DV) for potassium is 4700 mg.
   c. An automobile traveled 46.0 km on 1 gal of gasoline.
   d. The pesticide level in plums was 29 ppb.
   e. Silicon makes up 28.2% by mass of the Earth’s crust.

2.56 Write the equality and two conversion factors, and identify the numbers as exact or give the number of significant figures for each of the following:
   a. The Daily Value (DV) for iodine is 150 mcg.
   b. The nitrate level in well water was 32 ppm.

2.57 Write the equality and two conversion factors, and identify the numbers as exact or give the number of significant figures for each of the following:
   a. A calcium supplement contains 630 mg of calcium per tablet.
   b. The Daily Value (DV) for vitamin C is 60 mg.
   c. The label on a bottle reads 50 mg of atenolol per tablet.
   d. A low-dose aspirin contains 81 mg of aspirin per tablet.

2.58 Write the equality and two conversion factors, and identify the numbers as exact or give the number of significant figures for each of the following:
   a. The label on a bottle reads 10 mg of furosemide per 1 mL.
   b. The Daily Value (DV) for selenium is 70 mcg.
   c. An IV of normal saline solution has a flow rate of 85 mL per hour.
   d. One capsule of fish oil contains 360 mg of omega-3 fatty acids.

2.59 Write an equality and two conversion factors for each of the following medications in stock:
   a. 10 mg of Atarax per 5 mL of Atarax syrup
   b. 0.25 g of Lanoxin per 1 tablet of Lanoxin
   c. 300 mg of Motrin per 1 tablet of Motrin

2.60 Write an equality and two conversion factors for each of the following medications in stock:
   a. 2.5 mg of Coumadin per 1 tablet of Coumadin
   b. 100 mg of Clozapine per 1 tablet of Clozapine
   c. 1.5 g of Cefuroxime per 1 mL of Cefuroxime

2.6 Problem Solving Using Unit Conversion

LEARNING GOAL Use conversion factors to change from one unit to another.

The process of problem solving in chemistry often requires one or more conversion factors to change a given unit to the needed unit. For the problem, the unit of the given quantity and the unit of the needed quantity are identified. From there, the problem is set up with one or more conversion factors used to convert the given unit to the needed unit as seen in the following Sample Problem.

Given unit × one or more conversion factors = needed unit

SAMPLE PROBLEM 2.10 Problem Solving Using Conversion Factors

Greg’s doctor has ordered a PET scan of his heart. In radiological imaging such as PET or CT scans, dosages of pharmaceuticals are based on body mass. If Greg weighs 165 lb, what is his body mass in kilograms?
SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>165 lb</td>
<td>kilograms</td>
<td>conversion factor (kg/lb)</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to convert the given unit to the needed unit. The conversion factor relates the given unit in the U.S. system of measurement and the needed unit in the metric system.

pounds U.S.–Metric factor kilograms

STEP 3 State the equalities and conversion factors.

\[
\frac{1 \text{ kg}}{2.205 \text{ lb}} \quad \text{and} \quad \frac{2.205 \text{ lb}}{1 \text{ kg}}
\]

STEP 4 Set up the problem to cancel units and calculate the answer. Write the given, 165 lb, and multiply by the conversion factor that has the unit lb in the denominator (bottom number) to cancel out the given unit (lb) in the numerator.

\[
165 \text{ lb} \times \frac{1 \text{ kg}}{2.205 \text{ lb}} = 74.8 \text{ kg}
\]

Look at how the units cancel. The given unit lb cancels out and the needed unit kg is in the numerator. The unit you want in the final answer is the one that remains after all the other units have canceled out. This is a helpful way to check that you set up a problem properly.

\[
lb \times \frac{kg}{lb} = kg
\]

The calculator display gives the numerical answer, which is adjusted to give a final answer with the proper number of significant figures (SFs).

\[
165 \times 2.205 = 74.8 \text{ kg}
\]

The value of 74.8 combined with the unit, kg, gives the final answer of 74.8 kg. With few exceptions, answers to numerical problems contain a number and a unit.

STUDY CHECK 2.10

A total of 2500 mL of a boric acid antiseptic solution is prepared from boric acid concentrate. How many quarts of boric acid have been prepared?

ANSWER

2.6 qt
Using Two or More Conversion Factors

In problem solving, two or more conversion factors are often needed to complete the change of units. In setting up these problems, one factor follows the other. Each factor is arranged to cancel the preceding unit until the needed unit is obtained. Once the problem is set up to cancel units properly, the calculations can be done without writing intermediate results. The process is worth practicing until you understand unit cancellation, the steps on the calculator, and rounding off to give a final answer. In this text, when two or more conversion factors are required, the final answer will be based on obtaining a final calculator display and rounding off (or adding zeros) to give the correct number of significant figures.

SAMPLE PROBLEM 2.11 Using Two or More Conversion Factors

A doctor’s order for 0.50 g of Keflex is available as 250-mg tablets. How many tablets of Keflex are needed? In the following setup, fill in the missing parts of the conversion factors, show the canceled units, and give the correct answer:

\[
0.50 \text{ g Keflex} \times \frac{\text{mg Keflex}}{1 \text{ g Keflex}} \times \frac{1 \text{ tablet}}{\frac{250 \text{ mg Keflex}}{1}} = \text{tablets}
\]

TRY IT FIRST

SOLUTION

\[
0.50 \text{ g Keflex} \times \frac{1000 \text{ mg Keflex}}{1 \text{ g Keflex}} \times \frac{1 \text{ tablet}}{250 \text{ mg Keflex}} = 2 \text{ tablets}
\]

STUDY CHECK 2.11

A newborn has a length of 450 mm. In the following setup that calculates the baby’s length in inches, fill in the missing parts of the conversion factors, show the canceled units, and give the correct answer:

\[
450 \text{ mm} \times \frac{1 \text{ cm}}{\text{mm}} \times \frac{1 \text{ in.}}{\text{cm}} = \text{inches}
\]

ANSWER

\[
450 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} \times \frac{1 \text{ in.}}{2.54 \text{ cm}} = 18 \text{ in.}
\]

SAMPLE PROBLEM 2.12 Problem Solving Using Two Conversion Factors

During a volcanic eruption on Mauna Loa, the lava flowed at a rate of 33 m/min. At this rate, what distance, in kilometers, will the lava travel in 45 min?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>45 min</td>
<td>kilometers</td>
<td>conversion factors (m/min, km/m)</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to convert the given unit to the needed unit.

minutes Rate factor  meters Metric factor  kilometers
STEP 3 State the equalities and conversion factors. The rate of lava flow (33 m/min) is used as an equality as well as the metric equality for meters and kilometers. Then conversion factors can be written for both.

\[
\begin{align*}
1 \text{ min} &= 33 \text{ m} \\
33 \text{ m} &\quad \text{and} \quad 1 \text{ min} = 1000 \text{ m} \\
1 \text{ km} &= 1000 \text{ m} \quad \text{and} \quad 1 \text{ km} = 1000 \text{ m}
\end{align*}
\]

STEP 4 Set up the problem to cancel units and calculate the answer.

\[
45 \text{ min} \times \frac{33 \text{ m}}{1 \text{ min}} \times \frac{1 \text{ km}}{1000 \text{ m}}
\]

The calculation is done as follows:

\[
45 \times 33 \div 1000 = 1.485
\]

Calculator display

Because there are two significant figures in the measured quantities, we write the needed answer with two significant figures, 1.5, and the unit km to give the final answer, 1.5 km.

\[
\begin{align*}
45 \text{ min} &\quad \text{Two SFs} \\
33 \text{ m} &\quad \text{Exacting} \\
1 \text{ km} &\quad \text{Two SFs}
\end{align*}
\]

STUDY CHECK 2.12

How many hours are required for the lava in Sample Problem 2.12 to flow a distance of 5.0 km?

ANSWER

2.5 h

SAMPLE PROBLEM 2.13 Using a Percentage as a Conversion Factor

A person who exercises regularly has 16% body fat by mass. If this person weighs 155 lb, what is the mass, in kilograms, of body fat?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>155 lb body weight</td>
<td>kilograms of body fat</td>
<td>conversion factors (kg/lb, percent body fat)</td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to convert the given unit to the needed unit.

<table>
<thead>
<tr>
<th>pounds of body weight</th>
<th>U.S.–Metric factor</th>
<th>kilograms of body mass</th>
<th>Percentage factor</th>
<th>kilograms of body fat</th>
</tr>
</thead>
</table>

Exercising regularly helps reduce body fat.
STEP 3 State the equalities and conversion factors.

\[
\begin{align*}
1 \text{ kg of body mass} &= 2.205 \text{ lb of body weight} \\
&= \frac{1 \text{ kg body mass}}{1 \text{ kg body mass}} \times \frac{2.205 \text{ lb body weight}}{1 \text{ kg body mass}}
\end{align*}
\]

16 kg of body fat = 100 kg of body mass

\[
\frac{16 \text{ kg of body fat}}{100 \text{ kg body mass}} \times \frac{100 \text{ kg body mass}}{16 \text{ kg body fat}} = 1 \text{ kg of body fat}
\]

STEP 4 Set up the problem to cancel units and calculate the answer. We set up the problem using conversion factors to cancel each unit, starting with lb of body weight, until we obtain the final unit, kg of body fat, in the numerator.

\[
155 \text{ lb body weight} \times \frac{1 \text{ kg body mass}}{2.205 \text{ lb body weight}} \times \frac{16 \text{ kg body fat}}{100 \text{ kg body mass}} = 11 \text{ kg of body fat}
\]

STUDY CHECK 2.13

Uncooked lean ground beef can contain up to 22% fat by mass. How many grams of fat would be contained in 0.25 lb of the ground beef?

ANSWER

25 g of fat

CHEMISTRY LINK TO HEALTH

Toxicology and Risk–Benefit Assessment

Each day, we make choices about what we do or what we eat, often without thinking about the risks associated with these choices. We are aware of the risks of cancer from smoking or the risks of lead poisoning, and we know there is a greater risk of having an accident if we cross a street where there is no light or crosswalk.

A basic concept of toxicology is the statement of Paracelsus that the dose is the difference between a poison and a cure. To evaluate the level of danger from various substances, natural or synthetic, a risk assessment is made by exposing laboratory animals to the substances and monitoring the health effects. Often, doses very much greater than humans might ordinarily encounter are given to the test animals.

Many hazardous chemicals or substances have been identified by these tests. One measure of toxicity is the LD50, or lethal dose, which is the concentration of the substance that causes death in 50% of the test animals. A dosage is typically measured in milligrams per kilogram (mg/kg) of body mass or micrograms per kilogram (mcg/kg) of body mass.

Other evaluations need to be made, but it is easy to compare LD50 values. Parathion, a pesticide, with an LD50 of 3 mg/kg, would be highly toxic. This means that 3 mg of parathion per kg of body mass would be fatal to half the test animals. Table salt (sodium chloride) with an LD50 of 3300 mg/kg would have a much lower toxicity. You would need to ingest a huge amount of salt before any toxic effect would be observed. Although the risk to animals can be evaluated in the laboratory, it is more difficult to determine the impact in the environment since there is also a difference between continued exposure and a single, large dose of the substance.

TABLE 2.8 lists some LD50 values and compares substances in order of increasing toxicity.

<table>
<thead>
<tr>
<th>Substance</th>
<th>LD50 (mg/kg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Table sugar</td>
<td>29 700</td>
</tr>
<tr>
<td>Boric acid</td>
<td>5140</td>
</tr>
<tr>
<td>Baking soda</td>
<td>4220</td>
</tr>
<tr>
<td>Table salt</td>
<td>3300</td>
</tr>
<tr>
<td>Ethanol</td>
<td>2080</td>
</tr>
<tr>
<td>Aspirin</td>
<td>1100</td>
</tr>
<tr>
<td>Codeine</td>
<td>800</td>
</tr>
<tr>
<td>Oxytocodone</td>
<td>480</td>
</tr>
<tr>
<td>Caffeine</td>
<td>192</td>
</tr>
<tr>
<td>DDT</td>
<td>113</td>
</tr>
<tr>
<td>Cocaine (injected)</td>
<td>95</td>
</tr>
<tr>
<td>Dichlorvos (pesticide strips)</td>
<td>56</td>
</tr>
<tr>
<td>Ricin</td>
<td>30</td>
</tr>
<tr>
<td>Sodium cyanide</td>
<td>6</td>
</tr>
<tr>
<td>Parathion</td>
<td>3</td>
</tr>
</tbody>
</table>
2.6 Problem Solving Using Unit Conversion

LEARNING GOAL Use conversion factors to change from one unit to another.

2.61 When you convert one unit to another, how do you know which unit of the conversion factor to place in the denominator?

2.62 When you convert one unit to another, how do you know which unit of the conversion factor to place in the numerator?

2.63 Perform each of the following conversions using conversion factors:
   a. $4.42 \text{ mL to liters}$
   b. $8.65 \text{ m to nanometers}$
   c. $5.2 \times 10^8 \text{ g to megagrams}$
   d. $0.72 \text{ ks to milliseconds}$

2.64 Perform each of the following conversions using conversion factors:
   a. $4.82 \times 10^{-5} \text{ L to picoliters}$
   b. $575.2 \text{ dm to kilometers}$
   c. $5 \times 10^{-4} \text{ kg to micrograms}$
   d. $6.4 \times 10^{20} \text{ ps to seconds}$

2.65 Perform each of the following conversions using metric and U.S. conversion factors:
   a. $3.428 \text{ lb to kilograms}$
   b. $1.6 \text{ m to inches}$
   c. $4.2 \text{ L to quarts}$
   d. $0.672 \text{ ft to millimeters}$

2.66 Perform each of the following conversions using metric and U.S. conversion factors:
   a. $0.21 \text{ lb to grams}$
   b. $11.6 \text{ in. to centimeters}$
   c. $0.15 \text{ qt to milliliters}$
   d. $35.41 \text{ kg to pounds}$

2.67 Use metric conversion factors to solve each of the following problems:
   a. The height of a student is $175 \text{ cm}$. How tall is the student in meters?
   b. A cooler has a volume of $5000 \text{ mL}$. What is the capacity of the cooler in liters?
   c. A hummingbird has a mass of $0.0055 \text{ kg}$. What is the mass, in grams, of the hummingbird?
   d. A balloon has a volume of $3500 \text{ cm}^3$. What is the volume, in cubic meters?

2.68 Use metric conversion factors to solve each of the following problems:
   a. The Daily Value (DV) for phosphorus is $800 \text{ mg}$. How many grams of phosphorus are recommended?
   b. A glass of orange juice contains $3.2 \text{ dL}$ of juice. How many milliliters of orange juice are in the glass?
   c. A package of chocolate instant pudding contains $2840 \text{ mg}$ of sodium. How many grams of sodium are in the pudding?
   d. A park has an area of $150,000 \text{ m}^2$. What is the area, in square kilometers?

2.69 Solve each of the following problems using one or more conversion factors:
   a. A container holds $0.500 \text{ qt of liquid}$. How many milliliters of lemonade will it hold?
   b. What is the mass, in kilograms, of a person who weighs $175 \text{ lb}$?
   c. An athlete has $15\%$ body fat by mass. What is the weight of fat, in pounds, of a $74-\text{kg athlete}$?
   d. A plant fertilizer contains $15\%$ nitrogen (N) by mass. In a container of soluble plant food, there are $10.0 \text{ oz of fertilizer}$. How many grams of nitrogen are in the container?

2.70 Solve each of the following problems using one or more conversion factors:
   a. Wine is $12\%$ alcohol by volume. How many milliliters of alcohol are in $0.750 \text{ L}$ bottle of wine?
   b. Blueberry high-fiber muffins contain $51\%$ dietary fiber by mass. If a package with a net weight of $12 \text{ oz}$ contains six muffins, how many grams of fiber are in each muffin?
   c. A jar of crunchy peanut butter contains $1.43 \text{ kg of peanut butter}$. If you use $8.0\%$ of the peanut butter for a sandwich, how many ounces of peanut butter did you take out of the container?
   d. In a candy factory, the nutty chocolate bars contain $22.0\%$ pecans by mass. If $5.0 \text{ kg of pecans were used for candy last Tuesday}$, how many pounds of nutty chocolate bars were made?

Applications

2.71 Using conversion factors, solve each of the following clinical problems:
   a. You have used $250 \text{ L of distilled water}$ for a dialysis patient. How many gallons of water is that?
   b. A patient needs $0.024 \text{ g}$ of a sulfa drug. There are $8-\text{mg tablets in stock}$. How many tablets should be given?
   c. The daily dose of ampicillin for the treatment of an ear infection is $115 \text{ mg/kg of body weight}$. What is the daily dose for a $34-\text{lb child}$?
   d. You need $4.0 \text{ oz of a steroid ointment}$. How many grams of ointment does the pharmacist need to prepare?

2.72 Using conversion factors, solve each of the following clinical problems:
   a. The physician has ordered $1.0 \text{ g of tetracycline to be given}$ every six hours to a patient. If your stock on hand is $500-\text{mg tablets}$, how many will you need for one day’s treatment?
   b. An intramuscular medication is given at $5.00 \text{ mg/kg of body weight}$. What is the dose for a $180-\text{lb patient}$?
   c. A physician has ordered $0.50 \text{ mg of atropine, intramuscularly}$. If atropine were available as $0.10 \text{ mg/mL of solution}$, how many milliliters would you need to give?
   d. During surgery, a patient receives $5.0 \text{ pt of plasma}$. How many milliliters of plasma were given?

2.73 Using conversion factors, solve each of the following clinical problems:
   a. A nurse practitioner prepares $500. \text{ mL of an IV of normal saline solution}$ to be delivered at a rate of $80. \text{ mL/h}$. What is the infusion time, in hours, to deliver $500. \text{ mL}$?
   b. A nurse practitioner orders Medrol to be given $1.5 \text{ mg/kg of body weight}$. Medrol is an anti-inflammatory administered as an intramuscular injection. If a child weighs $72.6 \text{ lb}$ and the available stock of Medrol is $20. \text{ mg/mL}$, how many milliliters does the nurse administer to the child?
2.7 Density

**LEARNING GOAL** Calculate the density of a substance; use the density to calculate the mass or volume of a substance.

The mass and volume of any object can be measured. If we compare the mass of the object to its volume, we obtain a relationship called density.

\[
\text{Density} = \frac{\text{mass of substance}}{\text{volume of substance}}
\]

Every substance has a unique density, which distinguishes it from other substances. For example, lead has a density of 11.3 g/mL, whereas cork has a density of 0.26 g/mL. From these densities, we can predict if these substances will sink or float in water. If an object is less dense than a liquid, the object floats when placed in the liquid. If a substance, such as cork, is less dense than water, it will float. However, a lead object sinks because its density is greater than that of water (see FIGURE 2.10).

**FIGURE 2.10** Objects that sink in water are more dense than water; objects that float are less dense.

**ENGAGE** If a piece of iron sinks in water, how does its density compare to that of water?

Density is used in chemistry in many ways. For example, density can be used to identify a specific substance. If we calculate a density of a pure metal as 10.5 g/mL, then we could identify it as silver, but not gold or aluminum. Metals such as gold and silver have higher densities, whereas gases have low densities. In the metric system, the densities of solids and liquids are usually expressed as grams per cubic centimeter (g/cm\(^3\)) or grams per milliliter (g/mL). The densities of gases are usually stated as grams per liter (g/L). **TABLE 2.9** gives the densities of some common substances.

### Calculating Density

We can calculate the density of a substance from its mass and volume as shown in Sample Problem 2.14.

#### Table 2.8

<table>
<thead>
<tr>
<th>Substance</th>
<th>Density (g/mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cork</td>
<td>0.26</td>
</tr>
<tr>
<td>Ice</td>
<td>0.92</td>
</tr>
<tr>
<td>Water</td>
<td>1.00</td>
</tr>
<tr>
<td>Aluminum</td>
<td>2.70</td>
</tr>
<tr>
<td>Lead</td>
<td>11.3</td>
</tr>
</tbody>
</table>

#### Table 2.9

<table>
<thead>
<tr>
<th>Substance</th>
<th>Density (g/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Table salt</td>
<td>0.0001</td>
</tr>
<tr>
<td>Sodium cyanide</td>
<td>0.75</td>
</tr>
<tr>
<td>Aspirin</td>
<td>0.60</td>
</tr>
</tbody>
</table>

2.74 Using conversion factors, solve each of the following clinical problems:

a. A nurse practitioner prepares an injection of promethazine, an antihistamine used to treat allergic rhinitis. If the stock bottle is labeled 25 mg/mL and the order is a dose of 12.5 mg, how many milliliters will the nurse draw up in the syringe?

b. You are to give ampicillin 25 mg/kg to a child with a mass of 62 lb. If stock on hand is 250 mg/capsule, how many capsules should be given?

2.75 Using Table 2.8, calculate the number of grams that would provide the LD\(_{50}\) for a 175-lb person for each of the following:

a. Table salt
b. Sodium cyanide
c. Aspirin

2.76 Using Table 2.8, calculate the number of grams that would provide the LD\(_{50}\) for a 148-lb person for each of the following:

a. Ethanol
b. Ricin
c. Baking soda
SAMPLE PROBLEM 2.14 Calculating Density

High-density lipoprotein (HDL) is a type of cholesterol, sometimes called “good cholesterol,” that is measured in a routine blood test. If a 0.258-g sample of HDL has a volume of 0.215 cm³, what is the density of the HDL sample?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.258 g of HDL, 0.215 cm³</td>
<td>density (g/cm³) of HDL</td>
<td>density expression</td>
</tr>
</tbody>
</table>

STEP 2 Write the density expression.

\[
\text{Density} = \frac{\text{mass of substance}}{\text{volume of substance}}
\]

STEP 3 Express mass in grams and volume in cm³.

- Mass of HDL sample = 0.258 g
- Volume of HDL sample = 0.215 cm³

STEP 4 Substitute mass and volume into the density expression and calculate the density.

\[
\text{Density} = \frac{0.258 \text{ g}}{0.215 \text{ cm}^3} = 1.20 \text{ g/cm}^3 = 1.20 \text{ g/cm}^3
\]

Guide to Calculating Density

STEP 1 State the given and needed quantities.

STEP 2 Write the density expression.

STEP 3 Express mass in grams and volume in milliliters (mL) or cm³.

STEP 4 Substitute mass and volume into the density expression and calculate the density.
STUDY CHECK 2.14
Low-density lipoprotein (LDL), sometimes called “bad cholesterol,” is also measured in a routine blood test. If a 0.380-g sample of LDL has a volume of 0.362 cm³, what is the density of the LDL sample?

**ANSWER**
1.05 g/cm³

**Density of Solids Using Volume Displacement**
The volume of a solid can be determined by volume displacement. When a solid is completely submerged in water, it displaces a volume that is equal to the volume of the solid. In **FIGURE 2.11**, the water level rises from 35.5 mL to 45.0 mL after the zinc object is added. This means that 9.5 mL of water is displaced and that the volume of the object is 9.5 mL. The density of the zinc is calculated using volume displacement as follows:

\[
\text{Density} = \frac{68.60 \text{ g Zn}}{9.5 \text{ mL}} = 7.2 \text{ g/mL}
\]

**FIGURE 2.11** The density of a solid can be determined by volume displacement because a submerged object displaces a volume of water equal to its own volume.

Q: How is the volume of the zinc object determined?

**SAMPLE PROBLEM 2.15 Using Volume Displacement to Calculate Density**
A lead weight used in the belt of a scuba diver has a mass of 226 g. When the weight is carefully placed in a large graduated cylinder containing 200.0 cm³ of water, the water level rises to 220.0 cm³. What is the density (g/cm³) of the lead weight?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>226 g of lead, water level + lead = 220.0 cm³</td>
<td>density (g/cm³) of lead</td>
<td>density expression</td>
</tr>
<tr>
<td>water level (initial) = 200.0 cm³</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write the density expression.

\[
\text{Density} = \frac{\text{mass of substance}}{\text{volume of substance}}
\]
**STEP 3** Express mass in grams and volume in cm$^3$.

Mass of lead weight = 226 g

The volume of the lead weight is equal to the volume of water it displaced, which is calculated as follows:

Water level after object submerged = 220.0 cm$^3$

Water level before object submerged = −200.0 cm$^3$

Water displaced (volume of lead) = 20.0 cm$^3$

**STEP 4** Substitute mass and volume into the density expression and calculate the density. Be sure to use the volume of water displaced and *not* the initial volume of water.

Density $= \frac{226 \text{ g}}{20.0 \text{ cm}^3} = 11.3 \text{ g/cm}^3$

**STUDY CHECK 2.15**

A total of 0.500 lb of glass marbles is added to 425 mL of water. The water level rises to a volume of 528 mL. What is the density (g/cm$^3$) of the glass marbles?

**ANSWER**

2.20 g/cm$^3$

---

**CHEMISTRY LINK TO HEALTH**

**Bone Density**

The density of our bones determines their health and strength. Our bones are constantly gaining and losing minerals such as calcium, magnesium, and phosphate. In childhood, bones form at a faster rate than they break down. As we age, the breakdown of bone occurs more rapidly than new bone forms. As the loss of bone minerals increases, bones begin to thin, causing a decrease in mass and density. Thinner bones lack strength, which increases the risk of fracture. Hormonal changes, disease, and certain medications can also contribute to the thinning of bone. Eventually, a condition of severe thinning of bone known as *osteoporosis* may occur. *Scanning electron micrographs* (SEMs) show (a) normal bone and (b) bone with osteoporosis due to loss of bone minerals.

Bone density is often determined by passing low-dose X-rays through the narrow part at the top of the femur (hip) and the spine (c). These locations are where fractures are more likely to occur, especially as we age. Bones with high density will block more of the X-rays compared to bones that are less dense. The results of a bone density test are compared to a healthy young adult as well as to other people of the same age.

Recommendations to improve bone strength include supplements of calcium and vitamin D. Weight-bearing exercise such as walking and lifting weights can also improve muscle strength, which in turn increases bone strength.
Problem Solving Using Density

Density can be used as a conversion factor. For example, if the volume and the density of a sample are known, the mass in grams of the sample can be calculated as shown in Sample Problem 2.16.

**SAMPLE PROBLEM 2.16 Problem Solving with Density**

Greg has a blood volume of 5.9 qt. If the density of blood is 1.06 g/mL, what is the mass, in grams, of Greg’s blood?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>5.9 qt of blood</td>
<td>grams of blood</td>
<td>conversion factors (qt/mL, density of blood)</td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to calculate the needed quantity.

- quarts to U.S.–Metric factor
- milliliters to Density factor
- Density factor to grams

**STEP 3** Write the equalities and their conversion factors including density.

- \( 1 \text{ qt} = 946.4 \text{ mL} \)
- \( 1 \text{ mL of blood} = 1.06 \text{ g of blood} \)

**STEP 4** Set up the problem to calculate the needed quantity.

\[
7.5 \frac{\text{qt of blood}}{} \times \frac{946.4 \text{ mL}}{1 \text{ qt}} \times \frac{1.06 \text{ g blood}}{1 \text{ mL blood}} = 7500 \text{ g of blood}
\]

**STUDY CHECK 2.16**

During surgery, a patient receives 3.0 pt of blood. How many kilograms of blood (density = 1.06 g/mL) were needed for the transfusion?

**ANSWER**

1.5 kg of blood

**Specific Gravity**

Specific gravity (sp gr) is a relationship between the density of a substance and the density of water. Specific gravity is calculated by dividing the density of a sample by the density of water, which is 1.00 g/mL at 4 °C. A substance with a specific gravity of 1.00 has the same density as water (1.00 g/mL).

\[
\text{Specific gravity} = \frac{\text{density of sample}}{\text{density of water}}
\]

Specific gravity is one of the few unitless values you will encounter in chemistry. An instrument called a hydrometer is often used to measure the specific gravity of fluids such as wine or a sample of urine. The specific gravity of the urine helps evaluate the water balance.
in the body and the substances in the urine. In FIGURE 2.12, a hydrometer is used to measure the specific gravity of urine. The normal range of specific gravity for urine is 1.003 to 1.030. The specific gravity may decrease with type 2 diabetes and kidney disease. Increased specific gravity may occur with dehydration, kidney infection, and liver disease. In a clinic or hospital, a dipstick containing chemical pads is used to evaluate specific gravity.

FIGURE 2.12 A hydrometer is used to measure the specific gravity of urine, which, for adults, is 1.003 to 1.030.

If the hydrometer reading is 1.006, what is the density of the urine?

2.7 Density

Questions and Problems

2.7 Density

Learning Goal Calculate the density of a substance; use the density to calculate the mass or volume of a substance.

2.77 Determine the density (g/mL) for each of the following:
   a. A 20.0-mL sample of a salt solution that has a mass of 24.0 g.
   b. A cube of butter weighs 0.250 lb and has a volume of 130.3 mL.
   c. A gem has a mass of 4.50 g. When the gem is placed in a graduated cylinder containing 12.00 mL of water, the water level rises to 13.45 mL.
   d. A medication, if 3.00 mL has a mass of 3.85 g.

2.78 Determine the density (g/mL) for each of the following:
   a. The fluid in a car battery has a volume of 125 mL and a mass of 155 g.
   b. A plastic material weighs 2.68 lb and has a volume of 3.5 L.
   c. A 5.00-mL urine sample from a person suffering from diabetes mellitus has a mass of 5.025 g.
   d. A solid object with a mass of 1.65 lb and a volume of 170 mL.

2.79 What is the density (g/mL) of each of the following samples?
   a. A lightweight head on a golf club is made of titanium. The volume of a sample of titanium is 114 cm³ and the mass is 514.1 g.

b. A syrup is added to an empty container with a mass of 115.25 g. When 0.100 pt of syrup is added, the total mass of the container and syrup is 182.48 g.

c. A block of aluminum metal has a volume of 3.15 L and a mass of 8.51 kg.

2.80 What is the density (g/mL) of each of the following samples?
   a. An ebony carving has a mass of 275 g and a volume of 207 cm³.
   b. A 14.3-cm³ sample of tin has a mass of 0.104 kg.
   c. A bottle of acetone (fingernail polish remover) contains 55.0 mL of acetone with a mass of 43.5 g.

2.81 Use the density values in Table 2.9 to solve each of the following problems:
   a. How many liters of ethanol contain 1.50 kg of ethanol?
   b. How many grams of mercury are present in a barometer that holds 6.5 mL of mercury?
   c. A sculptor has prepared a mold for casting a silver figure. The figure has a volume of 225 cm³. How many ounces of silver are needed in the preparation of the silver figure?

2.82 Use the density values in Table 2.9 to solve each of the following problems:
   a. A graduated cylinder contains 18.0 mL of water. What is the new water level, in milliliters, after 35.6 g of silver metal is submerged in the water?
   b. A thermometer containing 8.3 g of mercury has broken. What volume, in cubic centimeters, of mercury spilled?
c. A fish tank holds 35 gal of water. How many kilograms of water are in the fish tank?

2.83 Use the density values in Table 2.9 to solve each of the following problems:
   a. What is the mass, in grams, of a cube of copper that has a volume of 74.1 cm³?
   b. How many kilograms of gasoline fill a 12.0-gal gas tank?
   c. What is the volume, in cubic centimeters, of an ice cube that has a mass of 27 g?

2.84 Use the density values in Table 2.9 to solve each of the following problems:
   a. If a bottle of olive oil contains 1.2 kg of olive oil, what is the volume, in milliliters, of the olive oil?
   b. A cannon ball made of iron has a volume of 115 cm³. What is the mass, in kilograms, of the cannon ball?
   c. A balloon filled with helium has a volume of 7.3 L. What is the mass, in grams, of helium in the balloon?

2.85 In an old trunk, you find a piece of metal that you think may be aluminum, silver, or lead. You take it to a lab, where you find it has a mass of 217 g and a volume of 19.2 cm³. Using Table 2.9, what is the metal you found?

2.86 Suppose you have two 100-mL graduated cylinders. In each cylinder, there is 40.0 mL of water. You also have two cubes: one is lead, and the other is aluminum. Each cube measures 2.0 cm on each side. After you carefully lower each cube into the water of its own cylinder, what will the new water level be in each of the cylinders?

Applications

2.87 Solve each of the following problems:
   a. A urine sample has a density of 1.030 g/mL. What is the specific gravity of the sample?
   b. A 20.0-mL sample of a glucose IV solution that has a mass of 20.6 g. What is the density of the glucose solution?
   c. The specific gravity of a vegetable oil is 0.92. What is the mass, in grams, of 750 mL of vegetable oil?
   d. A bottle containing 325 g of cleaning solution is used to clean hospital equipment. If the cleaning solution has a specific gravity of 0.850, what volume, in milliliters, of solution was used?

2.88 Solve each of the following problems:
   a. A glucose solution has a density of 1.02 g/mL. What is its specific gravity?
   b. A 0.200-mL sample of high-density lipoprotein (HDL) has a mass of 0.230 g. What is the density of the HDL?
   c. Butter has a specific gravity of 0.86. What is the mass, in grams, of 2.15 L of butter?
   d. A 5.000-mL urine sample has a mass of 5.025 g. If the normal range for the specific gravity of urine is 1.003 to 1.030, would the specific gravity of this urine sample indicate that the patient could have type 2 diabetes?

Applications

2.89 a. Write an equality and two conversion factors for Greg’s serum iron level.
   b. How many micrograms of iron were in the 8.0-mL sample of Greg’s blood?

2.90 a. Write an equality and two conversion factors for one tablet of the iron supplement.
   b. How many grams of iron will Greg consume in one week?
2.1 Units of Measurement

**LEARNING GOAL** Write the names and abbreviations for the metric or SI units used in measurements of length, volume, mass, temperature, and time.

- In science, physical quantities are described in units of the metric or International System of Units (SI).
- Some important units are meter (m) for length, liter (L) for volume, gram (g) and kilogram (kg) for mass, degree Celsius (°C) and kelvin (K) for temperature, and second (s) for time.

2.2 Measured Numbers and Significant Figures

**LEARNING GOAL** Identify a number as measured or exact; determine the number of significant figures in a measured number.

- A measured number is any number obtained by using a measuring device.
- An exact number is obtained by counting items or from a definition; no measuring device is needed.
- Significant figures are the numbers reported in a measurement including the estimated digit.
- Zeros in front of a decimal number or at the end of a nondecimal number are not significant.

2.3 Significant Figures in Calculations

**LEARNING GOAL** Adjust calculated answers to give the correct number of significant figures.

- In multiplication and division, the final answer is written so that it has the same number of significant figures as the measurement with the fewest significant figures.
- In addition and subtraction, the final answer is written so that it has the same number of decimal places as the measurement with the fewest decimal places.

2.4 Prefixes and Equalities

**LEARNING GOAL** Use the numerical values of prefixes to write a metric equality.

- A prefix placed in front of a metric or SI unit changes the size of the unit by factors of 10.
- Prefixes such as centi, milli, and micro provide smaller units; prefixes such as kilo, mega, and tera provide larger units.
- An equality shows the relationship between two units that measure the same quantity of length, volume, mass, or time.
- Examples of metric equalities are 1 m = 100 cm, 1 L = 1000 mL, 1 kg = 1000 g, and 1 min = 60 s.
2.5 Writing Conversion Factors

LEARNING GOAL Write a conversion factor for two units that describe the same quantity.

- Conversion factors are used to express a relationship in the form of a fraction.
- Two conversion factors can be written for any relationship in the metric or U.S. system.
- A percentage is written as a conversion factor by expressing matching units as the parts in 100 parts of the whole.
- Small ratios are expressed as parts per million (ppm) and parts per billion (ppb).

2.6 Problem Solving Using Unit Conversion

LEARNING GOAL Use conversion factors to change from one unit to another.

- Conversion factors are useful when changing a quantity expressed in one unit to a quantity expressed in another unit.

2.7 Density

LEARNING GOAL Calculate the density of a substance; use the density to calculate the mass or volume of a substance.

- The density of a substance is a ratio of its mass to its volume, usually g/mL or g/cm³.
- The units of density can be used to write conversion factors that convert between the mass and volume of a substance.
- Specific gravity (sp gr) compares the density of a substance to the density of water, 1.00 g/mL.

KEY TERMS

Celsius (°C) temperature scale A temperature scale on which water has a freezing point of 0 °C and a boiling point of 100 °C.
centimeter (cm) A unit of length in the metric system; there are 2.54 cm in 1 in.
conversion factor A ratio in which the numerator and denominator are quantities from an equality or given relationship. For example, the two conversion factors for the equality 1 kg = 2.205 lb are written as

\[
\frac{2.205 \text{ lb}}{1 \text{ kg}} \quad \text{and} \quad \frac{1 \text{ kg}}{2.205 \text{ lb}}
\]
cubic centimeter (cm³, cc) The volume of a cube that has 1-cm sides; 1 cm³ is equal to 1 mL.
cubic meter (m³) The volume of a cube that has sides that measure 1 m in length.
density The relationship of the mass of an object to its volume expressed as grams per cubic centimeter (g/cm³), grams per milliliter (g/mL), or grams per liter (g/L).
equality A relationship between two units that measure the same quantity.
exact number A number obtained by counting or by definition.
gram (g) The metric unit used in measurements of mass.
International System of Units (SI) The official system of measurement throughout the world except for the United States that modifies the metric system.
Kelvin (K) temperature scale A temperature scale on which the lowest possible temperature is 0 K.

kilogram (kg) A metric mass of 1000 g, equal to 2.205 lb. The kilogram is the SI standard unit of mass.
liter (L) The metric unit for volume that is slightly larger than a quart.
mass A measure of the quantity of material in an object.
measured number A number obtained when a quantity is determined by using a measuring device.
meter (m) The metric unit for length that is slightly longer than a yard. The meter is the SI standard unit of length.
metric system A system of measurement used by scientists and in most countries of the world.
milliliter (mL) A metric unit of volume equal to one-thousandth of a liter (0.001 L).
prefix The part of the name of a metric unit that precedes the base unit and specifies the size of the measurement. All prefixes are related on a decimal scale.
second (s) A unit of time used in both the SI and metric systems.
SI See International System of Units (SI).
significant figures (SFs) The numbers recorded in a measurement.
specific gravity (sp gr) A relationship between the density of a substance and the density of water:

\[ \text{sp gr} = \frac{\text{density of sample}}{\text{density of water}} \]
temperature An indicator of the hotness or coldness of an object.
volume (V) The amount of space occupied by a substance.
Rounding Off (2.3)
Calculator displays are rounded off to give the correct number of significant figures.
- If the first digit to be dropped is 4 or less, then it and all following digits are simply dropped from the number.
- If the first digit to be dropped is 5 or greater, then the last retained digit of the number is increased by 1.

Example: Round off each of the following to three significant figures:

- a. 3.60892 L
- b. 0.003870298 m
- c. 6 g

Answer:
- a. 3.61 L
- b. 0.00387 m
- c. 6.00 g

Counting Significant Figures (2.2)
The significant figures (SFs) are all the measured numbers including the last, estimated digit:

- All nonzero digits
- Zeros between nonzero digits
- Zeros within a decimal number
- All digits in a coefficient of a number written in scientific notation

An exact number is obtained from counting or a definition and has no effect on the number of significant figures in the final answer.

Example: State the number of significant figures in each of the following:

- a. 0.003045 mm
- b. 15,000 m
- c. 45.067 kg
- d. $5.30 \times 10^3$ g
- e. 2 cans of soda

Answer:
- a. four SFs
- b. two SFs
- c. five SFs
- d. three SFs
- e. exact

Using Significant Figures in Calculations (2.3)

- In multiplication or division, the final answer is written so that it has the same number of significant figures as the measurement with the fewest SFs.
- In addition or subtraction, the final answer is written so that it has the same number of decimal places as the measurement having the fewest decimal places.

Example: Perform the following calculations using measured numbers and give answers with the correct number of SFs:

- a. $4.05 \text{ m} \times 0.6078 \text{ m}$
- b. $450 \text{ g}$
- c. $0.758 \text{ g} + 3.10 \text{ g}$
- d. $13.538 \text{ km} - 8.6 \text{ km}$

Answer:
- a. $2.46 \text{ m}^2$
- b. $1.38 \text{ g/mL}$
- c. $3.86 \text{ g}$
- d. $4.9 \text{ km}$

Using Prefixes (2.4)

- In the metric and SI systems of units, a prefix attaches to any unit increases or decreases the size by some factor of 10.
- When the prefix centi is used with the unit meter, it becomes centimeter, a length that is one-hundredth of a meter (0.01 m).
- When the prefix milli is used with the unit meter, it becomes millimeter, a length that is one-thousandth of a meter (0.001 m).

Example: Complete the following statements with the correct prefix symbol:

- a. $1000 \text{ m} = 1 \ ____ \text{ m}$
- b. $0.01 \text{ g} = 1 \ ____ \text{ g}$

Answer:
- a. $1 \text{ km}$
- b. $1 \text{ cg}$

Writing Conversion Factors from Equalities (2.5)

- A conversion factor allows you to change from one unit to another.
- Two conversion factors can be written for any equality in the metric, U.S., or metric–U.S. systems of measurement.
- Two conversion factors can be written for a relationship stated within a problem.

Example: Write two conversion factors for the equality:

$$1 \text{ L} = 1000 \text{ mL}$$

Answer:

- $\frac{1000 \text{ mL}}{1 \text{ L}}$ and $\frac{1 \text{ L}}{1000 \text{ mL}}$

Using Conversion Factors (2.6)

In problem solving, conversion factors are used to cancel the given unit and to provide the needed unit for the answer.

- State the given and needed quantities.
- Write a plan to convert the given unit to the needed unit.
- State the equalities and conversion factors.
- Set up the problem to cancel units and calculate the answer.

Example: A computer chip has a width of 0.75 in. What is that distance in millimeters?

Answer:

$$0.75 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} \times \frac{10 \text{ mm}}{1 \text{ cm}} = 19 \text{ mm}$$

Using Density as a Conversion Factor (2.7)

Density is an equality of mass and volume for a substance, which is written as the density expression.

$$\text{Density} = \frac{\text{mass of substance}}{\text{volume of substance}}$$

Density is useful as a conversion factor to convert between mass and volume.

Example: The element tungsten used in light bulb filaments has a density of 19.3 g/cm$^3$. What is the volume, in cubic centimeters, of 250 g of tungsten?

Answer:

$$250 \text{ g} \times \frac{1 \text{ cm}^3}{19.3 \text{ g}} = 13 \text{ cm}^3$$
The chapter sections to review are shown in parentheses at the end of each question.

2.91 In which of the following pairs do both numbers contain the same number of significant figures? (2.2)
   a. 0.06890 m³ and 6.89 m³
   b. 0.42000 kg and 42 kg
   c. 10 m/s and 100 m/s
   d. 3.14 L and 3.14 × 10⁻³ L

2.92 How many significant figures are there in the following numbers? (2.2)
   a. 0.0109 g
   b. 250.000 cm³
   c. 0.09370 km
   d. 2.9838 × 10⁶ m

2.93 Indicate if each of the following is answered with an exact number or a measured number: (2.2)
   a. number of legs
   b. height of table
   c. number of chairs at the table
   d. area of tabletop

2.94 Measure the length of each of the objects in diagrams (a), (b), and (e) using the metric ruler in the figure. Indicate the number of significant figures for each and the estimated digit for each. (2.2)

2.95 State the temperature on the Celsius thermometer A to the correct number of significant figures: (2.3)

2.96 State the temperature on the Celsius thermometer B to the correct number of significant figures: (2.3)

2.97 The length of this rug is 38.4 in. and the width is 24.2 in. (2.3)
   a. What is the length of this rug, in centimeters?
   b. What is the width of this rug, in centimeters?
   c. How many significant figures are in the length measurement?
   d. Calculate the area of the rug, in square centimeters, to the correct number of significant figures.

2.98 A shipping box has a length of 7.00 in., a width of 6.00 in., and a height of 4.00 in. (2.3)
   a. What is the length of the box, in centimeters?
   b. What is the width of the box, in centimeters?
   c. How many significant figures are in the width measurement?
   d. Calculate the volume of the box, in cubic centimeters, to the correct number of significant figures.

2.99 Each of the following diagrams represents a container of water and a cube. Some cubes float while others sink. Match diagrams 1, 2, 3, or 4 with one of the following descriptions and explain your choices: (2.7)

   a. The cube has a greater density than water.
   b. The cube has a density that is 0.80 g/mL.
   c. The cube has a density that is one-half the density of water.
   d. The cube has the same density as water.
2.100 What is the density of the solid object that is weighed and submerged in water? (2.7)

2.101 Consider the following solids. The solids A, B, and C represent iron (D = 7.87 g/cm³), platinum (D = 21.5 g/cm³), and titanium (D = 4.51 g/cm³). If each has a mass of 10.0 g, what is the identity of each solid? (2.7)

2.102 A graduated cylinder contains three liquids A, B, and C, which have different densities and do not mix: dichloromethane (D = 1.33 g/mL), diethyl ether (D = 0.713 g/mL), and water (D = 1.00 g/mL). Identify the liquids A, B, and C in the cylinder. (2.7)

2.103 The gray cube has a density of 4.5 g/cm³. Is the density of the green cube the same, lower than, or higher than that of the gray cube? (2.7)

ADDITIONAL QUESTIONS AND PROBLEMS

2.105 Round off or add zeros to the following calculated answers to give a final answer with three significant figures: (2.2)
   a. 0.001 032 L  
   b. 6.8274 kg  
   c. 23.04 m  
   d. 32.965 s

2.106 Round off or add zeros to the following calculated answers to give a final answer with three significant figures: (2.2)
   a. 924.0924 m/s  
   b. 0.089 734 1 kg  
   c. 2.0924 × 10⁻³ g  
   d. 0.000 682 6 s

2.107 A dessert contains 137.25 g of vanilla ice cream, 84 g of fudge sauce, and 43.7 g of nuts. (2.3, 2.6)
   a. What is the total mass, in grams, of the dessert?  
   b. What is the total weight, in pounds, of the dessert?

2.108 A fish company delivers 22 kg of salmon, 5.5 kg of crab, and 3.48 kg of oysters to your seafood restaurant. (2.3, 2.6)
   a. What is the total mass, in kilograms, of the seafood?  
   b. What is the total number of pounds?

2.109 In France, grapes are 1.95 euros per kilogram. What is the cost of grapes, in dollars per pound, if the exchange rate is 1.14 dollars/euro? (2.6)

2.110 In Mexico, avocados are 48 pesos per kilogram. What is the cost, in cents, of an avocado that weighs 0.45 lb if the exchange rate is 15 pesos to the dollar? (2.6)

2.111 Bill’s recipe for onion soup calls for 4.0 lb of thinly sliced onions. If an onion has an average mass of 115 g, how many onions does Bill need? (2.6)

2.112 The price of 1 lb of potatoes is $1.75. If all the potatoes sold today at the store bring in $1420, how many kilograms of potatoes did grocery shoppers buy? (2.6)

2.113 During a workout at the gym, you set the treadmill at a pace of 55.0 m/min. How many minutes will you walk if you cover a distance of 7500 ft? (2.6)

2.114 The distance between two cities is 1700 km. How long will it take, in hours, to drive from one city to the other if your average speed is 63 mi/h? (2.6)

2.115 The water level in a graduated cylinder initially at 215 mL rises to 285 mL after a piece of lead is submerged. What is the mass, in grams, of the lead (see Table 2.9)? (2.7)

2.116 A graduated cylinder contains 160 mL of water. A 17.0-g piece of copper (density = 8.92 g/mL) and a 25.0-g piece of silver (density = 10.5 g/mL) are added. What is the new water level, in milliliters, in the cylinder? (2.7)

2.117 How many cubic centimeters (cm³) of mercury have the same mass as 2.0 L of ethanol (see Table 2.9)? (2.7)

2.118 What is the volume, in quarts, of 1.068 kg of olive oil? The density of olive oil is 0.92 g/mL. (2.7)
Applications

2.119 The following nutrition information is listed on a box of crackers: (2.6)
Serving size 0.50 oz (6 crackers)
Fat 4 g per serving; Sodium 140 mg per serving
a. If the box has a net weight (contents only) of 8.0 oz, about how many crackers are in the box?
b. If you ate 10 crackers, how many ounces of fat did you consume?
c. How many servings of crackers in part a would it take to obtain the Daily Value (DV) for sodium, which is 2400 mg?

2.120 The speed limit on a highway is 80 km/h. What is the speed in mi/h? (2.6)

2.121 To prevent bacterial infection, a doctor orders 4 tablets of amoxicillin per day for 10 days. If each tablet contains 250 mg of amoxicillin, how many ounces of the medication are given in 10 days? (2.6)

2.122 Celeste’s diet restricts her intake of protein to 24 g per day. If she eats 1.2 oz of protein, has she exceeded her protein limit for the day? (2.6)

2.123 The maximum adult dose of aspirin is 4000 mg in a 24-hour period. If a man has taken 4 pills in a day and each pill contains 0.8 g of aspirin, does it exceed the maximum dose? (2.6)

2.124 A doctor orders 15.0 mg of codeine. The cough syrup contains codeine at a concentration of 3. mg per 1. mL. How many mL of cough syrup should be administered to the patient? (2.6)

2.125 A balance measures mass to 0.001 g. If you determine the mass of an object that weighs about 31 g would you record the mass as 31 g, 31.1 g, 31.08 g, 31.075 g, or 31.0750? Explain your choice by writing two to three complete sentences that describe your thinking. (2.3)

2.126 When three students use the same meterstick to measure the length of a paper clip, they obtain results of 5.8 cm, 5.75 cm, and 5.76 cm. If the meterstick has millimeter markings, what are some reasons for the different values? (2.3)

2.127 The fuel consumption of a car is 1.8 gallons per 100 mile. If the cost of fuel per liter is $16.7, what is the fuel cost for travelling 55.0 km? (2.6)

2.128 A sunscreen contains 6% oxybenzone and 3% avobenzone by mass. If a bottle contains 40. g of sunscreen, how many kilograms of oxybenzone and avobenzone are needed to manufacture 1 ton of sunscreen? How many bottles of sunscreen can be produced for 1 ton of sunscreen? (2.6)

2.129 The volume of ice in the Arctic Ocean has decreased from 15 000 km³ in 1980 to 4500 km³ in 2012. What is the volume of water generated from the melted ice (Density of water at 0 °C is 0.9998 g/mL and density of ice at 0 °C is 0.92 g/mL)? (2.7)

2.130 A 30.0-g aluminum sphere and a 40.0-g iron sphere are both added to 65.7 mL of water contained in a graduated cylinder. What is the new water level, in milliliters, in the cylinder? (2.7)

2.131 In the manufacturing of computer chips, cylinders of silicon are cut into thin wafers that are 3.00 in. in diameter and have a mass of 1.50 g of silicon. How thick, in millimeters, is each wafer if silicon has a density of 2.33 g/cm³? (The volume of a cylinder is \( V = \pi r^2 h \).) (2.6, 2.7)

2.132 A package of aluminum foil is 66.7 yd long, 12 in. wide, and 0.000 30 in. thick. If aluminum has a density of 2.70 g/cm³, what is the mass, in grams, of the foil? (2.6, 2.7)

Applications

2.133 For a 180-lb person, calculate the quantity of each of the following that must be ingested to provide the LD₅₀ for caffeine given in Table 2.8: (2.6)

\[
\text{a. cups of coffee if one cup is 12 fl oz and there is 100. mg of caffeine per 6 fl oz of drip-brewed coffee}
\]

\[
\text{b. cans of cola if one can contains 50. mg of caffeine}
\]

\[
\text{c. tablets of NoDoz if one tablet contains 200. mg of caffeine}
\]

2.134 The food label of a bag of potato chips listed out the following nutritional values per 100 g of potato chips. If the serving size of potato chips is 1. oz, how many milligrams of each component would that person consume per serving? (2.6)

\[
\begin{align*}
\text{a. total fat} & \quad 33 \text{ g} \\
\text{sodium} & \quad 642 \text{ mg} \\
\text{total carbohydrate} & \quad 53.6 \text{ g} \\
\text{protein} & \quad 7.2 \text{ g} \\
\text{dietary fiber} & \quad 3.6 \text{ g}
\end{align*}
\]

2.135 a. Some athletes have as little as 3.0% body fat. If such a person has a body mass of 65 kg, how many pounds of body fat does that person have? (2.6)

b. In liposuction, a doctor removes fat deposits from a person’s body. If body fat has a density of 0.94 g/mL and 3.0 L of fat is removed, how many pounds of fat were removed from the patient?

2.136 A mouthwash is 21.6% ethanol by mass. If each bottle contains 0.358 pt of mouthwash with a density of 0.876 g/mL, how many kilograms of ethanol are in 180 bottles of the mouthwash? (2.6, 2.7)

A mouthwash may contain over 20% ethanol.
ANSWERS

Answers to Selected Questions and Problems

2.21 a. 5.0 × 10^4 L  b. 3.0 × 10^4 g  c. 1.0 × 10^3 m  d. 2.5 × 10^{-4} cm

2.23 a. 4  b. 2  c. 3

2.25 A calculator often gives more digits than the number of significant figures allowed in the answer.

2.27 a. 1.85 kg  b. 88.2 L  c. 0.00474 cm  d. 8810 m  e. 1.83 × 10^{3} s

2.29 a. 56.9 m  b. 0.00228 g  c. 11 500 s (1.15 × 10^{4} s)  d. 8.10 L

2.31 a. 1.6  b. 0.01  c. 27.6  d. 3.5  e. 0.14 (1.4 × 10^{-1})  f. 0.8 (8 × 10^{-1})

2.33 a. 53.54 cm  b. 127.6 g  c. 121.5 mL  d. 0.50 L

2.35 km/h is kilometers per hour; mi/h is miles per hour.

2.37 a. mg  b. dl  c. km  d. fg

2.39 a. centiliter  b. kilogram  c. millisecond  d. gigameter

2.41 a. 0.01  b. 10^{12}  c. 0.001  d. 0.1

2.43 a. decigram  b. microgram  c. kilogram  d. centigram

2.45 a. 100 cm  b. 1 × 10^9 nm  c. 0.001 m  d. 1000 mL

2.47 a. kilogram  b. milliliter  c. km  d. kl  e. nanometer

2.49 A conversion factor can be inverted to give a second conversion factor.

2.51 a. 1 m = 100 cm; \( \frac{100 \text{ cm}}{1 \text{ m}} \) and \( \frac{1 \text{ m}}{100 \text{ cm}} \)

b. 1 g = 1 × 10^9 ng; \( \frac{1 \times 10^9 \text{ ng}}{1 \text{ g}} \) and \( \frac{1 \text{ g}}{1 \times 10^9 \text{ ng}} \)

c. 1 kL = 1000 L; \( \frac{1000 \text{ L}}{1 \text{ kL}} \) and \( \frac{1 \text{ kL}}{1000 \text{ L}} \)

d. 1 s = 1000 ms; \( \frac{1000 \text{ ms}}{1 \text{ s}} \) and \( \frac{1 \text{ s}}{1000 \text{ ms}} \)

e. (1 m)^3 = (100 cm)^3; \( \frac{(100 \text{ cm})^3}{(1 \text{ m})^3} \) and \( \frac{(1 \text{ m})^3}{(100 \text{ cm})^3} \)

2.53 a. 1 yd = 3 ft; \( \frac{3 \text{ ft}}{1 \text{ yd}} \) and \( \frac{1 \text{ yd}}{3 \text{ ft}} \)

b. 1 kg = 2.205 lb; \( \frac{2.205 \text{ lb}}{1 \text{ kg}} \) and \( \frac{1 \text{ kg}}{2.205 \text{ lb}} \)

The 2.205 lb is measured: It has four SFs. The 1 kg is exact.

c. 1 min = 60 s; \( \frac{60 \text{ s}}{1 \text{ min}} \) and \( \frac{1 \text{ min}}{60 \text{ s}} \)

The 1 min and 60 s are both exact.

d. 1 gal = 27 mi; \( \frac{27 \text{ mi}}{1 \text{ gal}} \) and \( \frac{1 \text{ gal}}{27 \text{ mi}} \)

The 27 mi is measured: It has two SFs. The 1 gal is exact.

e. 93 g of silver = 100 g of sterling;

93 g silver \( \) and \( \frac{100 \text{ g sterling}}{93 \text{ g silver}} \)

The 93 g is measured: It has two SFs. The 100 g is exact.

2.55 a. 3.5 m = 1 s; \( \frac{3.5 \text{ m}}{1 \text{ s}} \) and \( \frac{1 \text{ s}}{3.5 \text{ m}} \)

The 3.5 m is measured: It has two SFs. The 1 s is exact.

b. 4700 mg of potassium = 1 day;

4700 mg potassium \( \) and \( \frac{1 \text{ day}}{4700 \text{ mg potassium}} \)

The 4700 mg is measured: It has two SFs. The 1 day is exact.

c. 46.0 km = 1 gal; \( \frac{46.0 \text{ km}}{1 \text{ gal}} \) and \( \frac{1 \text{ gal}}{46.0 \text{ km}} \)

The 46.0 km is measured: It has three SFs. The 1 gal is exact.

d. 29 mcg of pesticide = 1 kg of plums;

29 mcg pesticide \( \) and \( \frac{1 \text{ kg plums}}{29 \text{ mcg pesticide}} \)

The 29 mcg is measured: It has two SFs. The 1 kg is exact.

e. 28.2 g of silicon = 100 g of crust;

28.2 g silicon \( \) and \( \frac{100 \text{ g crust}}{28.2 \text{ g silicon}} \)

The 28.2 g is measured: It has three SFs. The 100 g is exact.

2.57 a. 1 tablet = 630 mg of calcium;

630 mg calcium \( \) and \( \frac{1 \text{ tablet}}{630 \text{ mg calcium}} \)

The 630 mg is measured: It has two SFs. The 1 tablet is exact.

b. 60 mg of vitamin C = 1 day;

60 mg vitamin C \( \) and \( \frac{1 \text{ day}}{60 \text{ mg vitamin C}} \)

The 60 mg is measured: It has one SF. The 1 day is exact.

c. 1 tablet = 50 mg of atenolol;

50 mg atenolol \( \) and \( \frac{1 \text{ tablet}}{50 \text{ mg atenolol}} \)

The 50 mg is measured: It has one SF. The 1 tablet is exact.
d. 1 tablet = 81 mg of aspirin;  
\[
\frac{81 \text{ mg aspirin}}{1 \text{ tablet}} \quad \text{and} \quad \frac{81 \text{ mg aspirin}}{1 \text{ tablet}}
\]

The 81 mg is measured: It has two SFs. The 1 tablet is exact.

2.59 a. 5 mL of syrup = 10 mg of Atarax;  
\[
\frac{10 \text{ mg Atarax}}{5 \text{ mL syrup}} \quad \text{and} \quad \frac{10 \text{ mg Atarax}}{5 \text{ mL syrup}}
\]
b. 1 tablet = 0.25 g of Lanoxin;  
\[
\frac{0.25 \text{ g Lanoxin}}{1 \text{ tablet}} \quad \text{and} \quad \frac{0.25 \text{ g Lanoxin}}{1 \text{ tablet}}
\]
c. 1 tablet = 300 mg of Motrin;  
\[
\frac{300 \text{ mg Motrin}}{1 \text{ tablet}} \quad \text{and} \quad \frac{300 \text{ mg Motrin}}{1 \text{ tablet}}
\]

2.61 The unit in the denominator must cancel with the preceding unit in the numerator.

2.63 a. 0.0442 L  
b. 8.65 × 10^9 nm  
c. 5.2 × 10^2 Mg  
d. 7.2 × 10^5 ms

2.65 a. 1.555 kg  
b. 63 in.  
c. 4.4 qt  
d. 205 mm

2.67 a. 1.75 m  
b. 5 L  
c. 5.5 g  
d. 3.5 × 10^{-3} m^3

2.69 a. 473 mL  
b. 79.4 kg  
c. 24 lb  
d. 43 g

2.71 a. 66 gal  
b. 3 tablets  
c. 1800 mg  
d. 110 g

2.73 a. 6.3 h  
b. 2.5 mL

2.75 a. 260 g  
b. 0.5 g  
c. 88 g

2.77 a. 1.20 g/mL  
b. 0.870 g/mL  
c. 3.10 g/mL  
d. 1.28 g/mL

2.79 a. 4.51 g/mL  
b. 1.42 g/mL  
c. 2.70 g/mL

2.81 a. 1.91 L of ethanol  
b. 88 g of mercury  
c. 83.3 oz of silver

2.83 a. 661 g  
b. 34 kg  
c. 29 cm^3

2.85 Since we calculate the density to be 11.3 g/cm^3, we identify the metal as lead.

2.87 a. 1.03  
b. 1.03 g/mL  
c. 690 g  
d. 382 mL

2.89 a. 42 mcg of iron = 1 dL of blood;  
\[
\frac{42 \text{ mcg iron}}{1 \text{ dL blood}} \quad \text{and} \quad \frac{1 \text{ dL blood}}{42 \text{ mcg iron}}
\]
b. 3.4 mcg of iron

2.91 c and d

2.93 a. exact  
b. measured  
c. exact  
d. measured

2.95 61.5 °C

2.97 a. 97.5 cm  
b. 61.5 cm  
c. three SFs  
d. 6.00 × 10^3 cm^2

2.99 a. Diagram 3; a cube that has a greater density than the water will sink to the bottom.  
b. Diagram 4; a cube with a density of 0.80 g/mL will be about two-thirds submerged in the water.  
c. Diagram 1; a cube with a density that is one-half the density of water will be one-half submerged in the water.  
d. Diagram 2; a cube with the same density as water will float just at the surface of the water.

2.101 A would be platinum with the highest density (21.5 g/cm^3) and the smallest volume.  
B would be iron with the intermediate density (7.87 g/cm^3) and the intermediate volume.  
C would be titanium with the lowest density (4.51 g/cm^3) and the largest volume.

2.103 The green cube has the same volume as the gray cube. However, the green cube has a larger mass on the scale, which means that its mass/volume ratio is larger. Thus, the density of the green cube is higher than the density of the gray cube.

2.105 a. 0.001 03 L (1.03 × 10^{-3} L)  
b. 6.83 kg  
c. 23.0 m  
d. 33.0 s

2.107 a. 265 g  
b. 0.584 lb

2.109 $1.01 per lb

2.111 16 onions

2.113 42 min

2.115 790 g

2.117 116 cm^3 of mercury

2.119 a. 96 crackers  
b. 0.2 oz of fat  
c. 17 servings

2.121 0.35 oz

2.123 The amount of aspirin intake = 4 × 800 = 3200. mg (4 SFs). It does not exceed the maximum dose.

2.125 You would record the mass as 31.075 g. Since the balance will weigh to the nearest 0.001 g, the mass value would be reported to 0.001 g.

2.127 $38.9 (3 SFs)

2.129 9662 km^3

2.131 0.141 mm

2.133 a. 78 cups  
b. 310 cans  
c. 80 tablets

2.135 a. 4.3 lb of body fat  
b. 6.2 lb
CHARLES IS 13 YEARS OLD and overweight. His doctor is worried that Charles is at risk for type 2 diabetes and advises his mother to make an appointment with a dietitian. Daniel, a dietitian, explains to them that choosing the appropriate foods is important to living a healthy lifestyle, losing weight, and preventing or managing diabetes.

Daniel also explains that food contains potential or stored energy and different foods contain different amounts of potential energy. For instance, carbohydrates contain 4 kcal/g (17 kJ/g) whereas fats contain 9 kcal/g (38 kJ/g). He then explains that diets high in fat require more exercise to burn the fats, as they contain more potential energy. When Daniel looks at Charles’ typical daily diet, he calculates that Charles is obtaining 3800 kcal for one day. The American Heart Association recommends 1800 kcal for boys 9 to 13 years of age. Daniel encourages Charles and his mother to include whole grains, fruits, and vegetables in their diet instead of foods high in fat. They also discuss food labels and the fact that smaller serving sizes of healthy foods are necessary to lose weight. Daniel also recommends that Charles participate in at least 60 min of exercise every day. Before leaving, Charles and his mother are given a menu for the following two weeks, as well as a diary to keep track of what, and how much, they actually consume.

CAREER
Dietitian

Dietitians specialize in helping individuals learn about good nutrition and the need for a balanced diet. This requires them to understand biochemical processes, the importance of vitamins, and food labels, as well as the differences between carbohydrates, fats, and proteins in terms of their energy value and how they are metabolized. Dietitians work in a variety of environments, including hospitals, nursing homes, school cafeterias, and public health clinics. In these environments, they create specialized diets for individuals diagnosed with a specific disease or create meal plans for those in a nursing home.
CHAPTER READINESS*

**KEY MATH SKILLS**  
- Using Positive and Negative Numbers in Calculations (1.4)  
- Solving Equations (1.4)  
- Interpreting Graphs (1.4)  
- Converting between Standard Numbers and Scientific Notation (1.5)  
- Rounding Off (2.3)  

**CORE CHEMISTRY SKILLS**  
- Counting Significant Figures (2.2)  
- Using Significant Figures in Calculations (2.3)  
- Writing Conversion Factors from Equivalences (2.5)  
- Using Conversion Factors (2.6)  

*These Key Math Skills and Core Chemistry Skills from previous chapters are listed here for your review as you proceed to the new material in this chapter.

**LOOKING AHEAD**

| 3.1 Classification of Matter | 3.3 Temperature | 3.5 Specific Heat | 3.2 States and Properties of Matter | 3.4 Energy | 3.6 Energy and Nutrition |

**CORE CHEMISTRY SKILL**  
Classifying Matter

**3.1 Classification of Matter**

**LEARNING GOAL** Classify examples of matter as pure substances or mixtures.

**Matter** is anything that has mass and occupies space. Matter is everywhere around us: the orange juice we had for breakfast, the water we put in the coffee maker, the plastic bag we put our sandwich in, our toothbrush and toothpaste, the oxygen we inhale, and the carbon dioxide we exhale. To a scientist, all of this material is matter. The different types of matter are classified by their composition.

**Pure Substances: Elements and Compounds**

A **pure substance** is matter that has a fixed or definite composition. There are two kinds of pure substances: elements and compounds. An **element**, the simplest type of a pure substance, is composed of only one type of material such as silver, iron, or aluminum. Every element is composed of **atoms**, which are extremely tiny particles that make up each type of matter. Silver is composed of silver atoms, iron of iron atoms, and aluminum of aluminum atoms. A full list of the elements is found on the inside front cover of this text.

A **compound** is also a pure substance, but it consists of atoms of two or more elements always chemically combined in the same proportion. In compounds, the atoms are held together by attractions called **bonds**, which form small groups of atoms called **molecules**. For example, a molecule of the compound water has two hydrogen atoms for every one oxygen atom and is represented by the formula H₂O. This means that water found anywhere always has the same composition of H₂O. Another compound that consists of a chemical combination of hydrogen and oxygen is hydrogen peroxide. It has two hydrogen atoms for every two oxygen atoms and is represented by the formula H₂O₂. Thus, water (H₂O) and hydrogen peroxide (H₂O₂) are different compounds that have different properties even though they contain the same elements, hydrogen and oxygen.

Pure substances that are compounds can be broken down by chemical processes into their elements. They cannot be broken down through physical methods such as boiling or sifting. For example, ordinary table salt consists of the compound NaCl, which can be separated by chemical processes into sodium metal and chlorine gas, as seen in **FIGURE 3.1**. Elements cannot be broken down further.
Mixtures

In a **mixture**, two or more different substances are physically mixed, but not chemically combined. Much of the matter in our everyday lives consists of mixtures. The air we breathe is a mixture of mostly oxygen and nitrogen gases. The steel in buildings and railroad tracks is a mixture of iron, nickel, carbon, and chromium. The brass in doorknobs and fixtures is a mixture of zinc and copper (see **Figure 3.2**). Tea, coffee, and ocean water are mixtures too. Unlike compounds, the proportions of substances in a mixture are not consistent but can vary. For example, two sugar–water mixtures may look the same, but the one with the higher ratio of sugar to water would taste sweeter.

**Figure 3.1** The decomposition of salt, NaCl, produces the elements sodium and chlorine.

How do elements and compounds differ?

**Figure 3.2** Matter is organized by its components: elements, compounds, and mixtures.

(a) The element copper consists of copper atoms. (b) The compound water consists of \(H_2O\) molecules. (c) Brass is a homogeneous mixture of copper and zinc atoms. (d) Copper metal in water is a heterogeneous mixture of copper atoms and \(H_2O\) molecules.

Why are copper and water pure substances, but brass is a mixture?
Physical processes can be used to separate mixtures because there are no chemical interactions between the components. For example, different coins, such as nickels, dimes, and quarters, can be separated by size; iron particles mixed with sand can be picked up with a magnet; and water is separated from cooked spaghetti by using a strainer (see FIGURE 3.3).

Types of Mixtures
Mixtures are classified further as homogeneous or heterogeneous. In a homogeneous mixture, also called a solution, the composition is uniform throughout the sample. Familiar examples of homogeneous mixtures are air, which contains oxygen and nitrogen gases, and seawater, a solution of salt and water.

In a heterogeneous mixture, the components do not have a uniform composition throughout the sample. For example, a mixture of oil and water is heterogeneous because the oil floats on the surface of the water. Other examples of heterogeneous mixtures are a cookie with raisins and orange juice with pulp.

In the chemistry laboratory, mixtures are separated by various methods. Solids are separated from liquids by filtration, which involves pouring a mixture through a filter paper set in a funnel. In chromatography, different components of a liquid mixture separate as they move at different rates up the surface of a piece of chromatography paper.

SAMPLE PROBLEM 3.1 Classifying Mixtures
Classify each of the following as a pure substance (element or compound) or a mixture (homogeneous or heterogeneous):

a. copper in copper wire
b. a chocolate-chip cookie
c. nitrox, a combination of oxygen and nitrogen used to fill scuba tanks

TRY IT FIRST

SOLUTION

a. Copper is an element, which is a pure substance.
b. A chocolate-chip cookie does not have a uniform composition, which makes it a heterogeneous mixture.
c. The gases oxygen and nitrogen have a uniform composition in nitrox, which makes it a homogeneous mixture.

STUDY CHECK 3.1

A salad dressing is prepared with oil, vinegar, and chunks of blue cheese. Is this a homogeneous or heterogeneous mixture?

ANSWER

heterogeneous mixture
The air we breathe is composed mostly of the gases oxygen (21%) and nitrogen (79%). The homogeneous breathing mixtures used by scuba divers differ from the air we breathe depending on the depth of the dive. Nitrox is a mixture of oxygen and nitrogen, but with more oxygen gas (up to 32%) and less nitrogen gas (68%) than air. A breathing mixture with less nitrogen gas decreases the risk of nitrogen narcosis associated with breathing regular air while diving. Heliox contains oxygen and helium, which is typically used for diving to more than 200 ft. By replacing nitrogen with helium, nitrogen narcosis does not occur. However, at dive depths over 300 ft, helium is associated with severe shaking and body temperature drop.

A breathing mixture used for dives over 400 ft is trimix, which contains oxygen, helium, and some nitrogen. The addition of some oxygen lessens the problem of shaking that comes with breathing high levels of helium. Heliox and trimix are used only by professional, military, or other highly trained divers.

In hospitals, heliox may be used as a treatment for respiratory disorders and lung constriction in adults and premature infants. Heliox is less dense than air, which reduces the effort of breathing and helps distribute the oxygen gas to the tissues.

A nitrox mixture is used to fill scuba tanks.

### QUESTIONS AND PROBLEMS

#### 3.1 Classification of Matter

**LEARNING GOAL** Classify examples of matter as pure substances or mixtures.

3.1 Classify each of the following as a pure substance or a mixture:
- a. baking soda (NaHCO₃)
- b. a blueberry muffin
- c. ice (H₂O)
- d. zinc (Zn)
- e. trimix (oxygen, nitrogen, and helium) in a scuba tank

3.2 Classify each of the following as a pure substance or a mixture:
- a. a soft drink
- b. propane (C₃H₈)
- c. a cheese sandwich
- d. an iron (Fe) nail
- e. salt substitute (KCl)

3.3 Classify each of the following pure substances as an element or a compound:
- a. silicon (Si)
- b. hydrogen peroxide (H₂O₂)
- c. oxygen (O₂)
- d. rust (Fe₂O₃)
- e. methane (CH₄) in natural gas

3.4 Classify each of the following pure substances as an element or a compound:
- a. helium gas (He)
- b. sulfur (S)
- c. sugar (C₁₂H₂₂O₁₁)
- d. mercury (Hg) in a thermometer

3.5 Classify each of the following mixtures as homogeneous or heterogeneous:
- a. vegetable soup
- b. seawater
- c. tea
- d. tea with ice and lemon slices
- e. fruit salad

3.6 Classify each of the following mixtures as homogeneous or heterogeneous:
- a. nonfat milk
- b. chocolate-chip ice cream
- c. gasoline
- d. peanut butter sandwich

#### 3.2 States and Properties of Matter

**LEARNING GOAL** Identify the states and the physical and chemical properties of matter.

On Earth, matter exists in one of three physical forms called the *states of matter*: solids, liquids, and gases. Water is a familiar example that we routinely observe in all three states. In the solid state, water can be an ice cube or a snowflake. It is a liquid when it comes out of a faucet or fills a pool. Water forms a gas, or vapor, when it evaporates from wet clothes or boils in a pan. A solid, such as a pebble or a baseball, has a definite shape and volume. You can probably recognize several solids within your reach right now such as books, pencils, or a computer mouse. In a solid, strong attractive forces hold the particles such as atoms or molecules close together. The particles in a solid are arranged in such a rigid pattern, their only movement is to vibrate slowly in fixed positions. For many solids, this rigid structure produces a crystal such as that seen in amethyst.

Amethyst, a solid, is a purple form of quartz (SiO₂).
A liquid has a definite volume, but not a definite shape. In a liquid, the particles move in random directions but are sufficiently attracted to each other to maintain a definite volume, although not a rigid structure. Thus, when water, oil, or vinegar is poured from one container to another, the liquid maintains its own volume but takes the shape of the new container.

A gas does not have a definite shape or volume. In a gas, the particles are far apart, have little attraction to each other, and move at high speeds, taking the shape and volume of their container. When you inflate a bicycle tire, the air, which is a gas, fills the entire volume of the tire. The propane gas in a tank fills the entire volume of the tank. TABLE 3.1 compares the three states of matter.

### Physical Properties and Physical Changes

One way to describe matter is to observe its properties. For example, if you were asked to describe yourself, you might list your characteristics such as the color of your eyes and skin or the length, color, and texture of your hair. **Physical properties** are those characteristics that can be observed or measured without affecting the identity of a substance. In chemistry, typical physical properties include the shape, color, melting point, boiling point, and physical state of a substance. For example, a penny has the physical properties of a round shape, an orange-red color, solid state, and a shiny luster. TABLE 3.2 gives more examples of physical properties of copper found in pennies, electrical wiring, and copper pans.

Water is a substance that is commonly found in all three states: solid, liquid, and gas. When matter undergoes a **physical change**, its state, size, or its appearance will change, but its composition remains the same. The solid state of water, snow or ice, has a different appearance than its liquid or gaseous state, but all three states are water.
The physical appearance of a substance can change in other ways too. Suppose that you dissolve some salt in water. The appearance of the salt changes, but you could re-form the salt crystals by evaporating the water. Thus, in a physical change of state, no new substances are produced.

**Chemical Properties and Chemical Changes**

**Chemical properties** are those that describe the ability of a substance to change into a new substance. When a **chemical change** takes place, the original substance is converted into one or more new substances, which have different physical and chemical properties. For example, the rusting or corrosion of a metal, such as iron, is a chemical property. In the rain, an iron (Fe) nail undergoes a chemical change when it reacts with oxygen (O₂) to form rust (Fe₂O₃). A chemical change has taken place: Rust is a new substance with new physical and chemical properties. **TABLE 3.3** gives examples of some physical and chemical changes. **TABLE 3.4** summarizes physical and chemical properties and changes.

---

**TABLE 3.3** Examples of Some Physical and Chemical Changes

<table>
<thead>
<tr>
<th>Physical Changes</th>
<th>Chemical Changes</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water boils to form water vapor.</td>
<td>Shiny, silver metal reacts in air to give a black, grainy coating.</td>
</tr>
<tr>
<td>Copper is drawn into thin copper wires.</td>
<td>A piece of wood burns with a bright flame and produces heat, ashes, carbon dioxide, and water vapor.</td>
</tr>
<tr>
<td>Sugar dissolves in water to form a solution.</td>
<td>Heating white, granular sugar forms a smooth, caramel-colored substance.</td>
</tr>
<tr>
<td>Paper is cut into tiny pieces of confetti.</td>
<td>Iron, which is gray and shiny, combines with oxygen to form orange-red rust.</td>
</tr>
</tbody>
</table>

**TABLE 3.4** Summary of Physical and Chemical Properties and Changes

<table>
<thead>
<tr>
<th>Physical Property</th>
<th>Chemical Property</th>
</tr>
</thead>
<tbody>
<tr>
<td>Physical</td>
<td>A characteristic of a substance: color, shape, odor, luster, size, melting point, or density.</td>
</tr>
<tr>
<td>Change</td>
<td>A change in a physical property that retains the identity of the substance: a change of state, a change in size, or a change in shape.</td>
</tr>
<tr>
<td>Chemical Property</td>
<td>A characteristic that indicates the ability of a substance to form another substance: paper can burn, iron can rust, silver can tarnish.</td>
</tr>
<tr>
<td>Change</td>
<td>A change in which the original substance is converted to one or more new substances: paper burns, iron rusts, silver tarnishes.</td>
</tr>
</tbody>
</table>
Ch 3
CHAPTER 3  Matter and Energy

A gold ingot is hammered to form gold leaf.

3.2 Physical and Chemical Changes

Classify each of the following as a physical or chemical change:

a. A gold ingot is hammered to form gold leaf.
b. Gasoline burns in air.
c. Garlic is chopped into small pieces.

TRY IT FIRST

SOLUTION

a. A physical change occurs when the gold ingot changes shape.
b. A chemical change occurs when gasoline burns and forms different substances with new properties.
c. A physical change occurs when the size of the garlic pieces changes.

STUDY CHECK 3.2

Classify each of the following as a physical or chemical change:

a. Water freezes on a pond.
b. Gas bubbles form when baking powder is placed in vinegar.
c. A log is cut for firewood.

ANSWER

a. physical change    b. chemical change    c. physical change

QUESTIONS AND PROBLEMS

3.2 States and Properties of Matter

LEARNING GOAL  Identify the states and the physical and chemical properties of matter.

3.7 Indicate whether each of the following describes a gas, a liquid, or a solid:

a. The breathing mixture in a scuba tank has no definite volume or shape.
b. The neon atoms in a lighting display do not interact with each other.
c. The particles in an ice cube are held in a rigid structure.

3.8 Indicate whether each of the following describes a gas, a liquid, or a solid:

a. Lemonade has a definite volume but takes the shape of its container.
b. The particles in a tank of oxygen are very far apart.
c. Helium occupies the entire volume of a balloon.

3.9 Describe each of the following as a physical or chemical property:

a. Chromium is a steel-gray solid.
b. Hydrogen reacts readily with oxygen.
c. A patient has a temperature of 40.2 °C.
d. Milk will sour when left in a warm room.
e. Butane gas in an igniter burns in oxygen.

3.10 Describe each of the following as a physical or chemical property:

a. Neon is a colorless gas at room temperature.
b. Apple slices turn brown when they are exposed to air.
c. Phosphorus will ignite when exposed to air.
d. At room temperature, mercury is a liquid.
e. Propane gas is compressed to a liquid for placement in a small cylinder.

3.11 What type of change, physical or chemical, takes place in each of the following?

a. Water vapor condenses to form rain.
b. Cesium metal reacts explosively with water.
c. Gold melts at 1064 °C.
d. A puzzle is cut into 1000 pieces.
e. Cheese is grated.

3.12 What type of change, physical or chemical, takes place in each of the following?

a. Pie dough is rolled into thin pieces for a crust.
b. A silver pin tarnishes in the air.
c. A tree is cut into boards at a saw mill.
d. Food is digested.
e. A chocolate bar melts.

3.13 Describe each property of the element fluorine as physical or chemical.

a. is highly reactive
b. is a gas at room temperature
c. has a pale, yellow color
d. will explode in the presence of hydrogen
e. has a melting point of −220 °C

3.14 Describe each property of the element zirconium as physical or chemical.

a. melts at 1852 °C
b. is resistant to corrosion
c. has a grayish white color
d. ignites spontaneously in air when finely divided
e. is a shiny metal
3.3 Temperature

LEARNING GOAL Given a temperature, calculate a corresponding temperature on another scale.

The small particles, atoms and molecules, in matter are constantly moving due to heat or thermal energy. At higher temperatures, they move faster; at lower temperatures, they move slower.

Temperatures in science are measured and reported in Celsius (°C) units. On the Celsius scale, the reference points are the freezing point of water, defined as 0 °C, and the boiling point, 100 °C. In the United States, everyday temperatures are commonly reported in Fahrenheit (°F) units. On the Fahrenheit scale, water freezes at 32 °F and boils at 212 °F. A typical room temperature of 22 °C would be the same as 72 °F. Normal human body temperature is 37.0 °C, which is the same temperature as 98.6 °F.

On the Celsius and Fahrenheit temperature scales, the temperature difference between freezing and boiling is divided into smaller units called degrees. On the Celsius scale, there are 100 degrees Celsius between the freezing and boiling points of water, whereas the Fahrenheit scale has 180 degrees Fahrenheit between the freezing and boiling points of water. That makes a degree Celsius almost twice the size of a degree Fahrenheit: 1 °C = 1.8 °F (see FIGURE 3.4).

\[
\frac{180 \text{ degrees Fahrenheit}}{100 \text{ degrees Celsius}} = 1.8 \text{ °F per °C}
\]

We can write a temperature equation that relates a Fahrenheit temperature and its corresponding Celsius temperature.

\[ T_F = 1.8(T_C) + 32 \]

Temperature equation to obtain degrees Fahrenheit

Changes °C to °F
Adjusts freezing point

\[ T_F = 1.8(37.0) + 32 = 98.6 \]

\[ T_C = 5(98.6 - 32) = 37.0 \]

\[ T_K = 373 \]

As shown in FIGURE 3.4, water freezes at 0 °C (32 °F) and boils at 100 °C (212 °F) on the Celsius and Fahrenheit scales respectively. The normal human body temperature is 37.0 °C (98.6 °F).

FIGURE 3.4 ▶ A comparison of the Fahrenheit, Celsius, and Kelvin temperature scales between the freezing and boiling points of water.

Q What is the difference in the freezing points of water on the Celsius and Fahrenheit temperature scales?

ENGAGE

Why is a degree Celsius a larger unit of temperature than a degree Fahrenheit?
In this equation, the Celsius temperature is multiplied by 1.8 to change °C to °F; then 32 is added to adjust the freezing point from 0 °C to the Fahrenheit freezing point, 32 °F. The values, 1.8 and 32, used in the temperature equation are exact numbers and are not used to determine significant figures in the answer.

To convert from degrees Fahrenheit to degrees Celsius, the temperature equation is rearranged to solve for \( T_C \). First, we subtract 32 from both sides since we must apply the same operation to both sides of the equation.

\[
T_F - 32 = 1.8(T_C) + 32 - 32
\]
\[
T_F - 32 = 1.8(T_C)
\]

Second, we solve the equation for \( T_C \) by dividing both sides by 1.8.

\[
\frac{T_F - 32}{1.8} = \frac{1.8(T_C)}{1.8}
\]
\[
\frac{T_F - 32}{1.8} = T_C \quad \text{Temperature equation to obtain degrees Celsius}
\]

Scientists have learned that the coldest temperature possible is \(-273 °C \) (more precisely, \(-273.15 °C \)), On the Kelvin scale, this temperature, called absolute zero, has the value of 0 K. Units on the Kelvin scale are called kelvins (K); no degree symbol is used. Because there are no lower temperatures, the Kelvin scale has no negative temperature values. Between the freezing point of water, 273 K, and the boiling point, 373 K, there are 100 kelvins, which makes a kelvin equal in size to a degree Celsius.

\[1 \text{ K} = 1 °C\]

We can write an equation that relates a Celsius temperature to its corresponding Kelvin temperature by adding 273 to the Celsius temperature. Table 3.5 gives a comparison of some temperatures on the three scales.

\[T_K = T_C + 273 \quad \text{Temperature equation to obtain kelvins}\]

An antifreeze mixture in a car radiator will not freeze until the temperature drops to \(-37 °C \). We can calculate the temperature of the antifreeze mixture in kelvins by adding 273 to the temperature in degrees Celsius.

\[T_K = -37 °C + 273 = 236 \text{ K}\]

### Table 3.5 A Comparison of Temperatures

<table>
<thead>
<tr>
<th>Example</th>
<th>Fahrenheit (°F)</th>
<th>Celsius (°C)</th>
<th>Kelvin (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sun</td>
<td>9937</td>
<td>5503</td>
<td>5776</td>
</tr>
<tr>
<td>A hot oven</td>
<td>450</td>
<td>232</td>
<td>505</td>
</tr>
<tr>
<td>Water boils</td>
<td>212</td>
<td>100</td>
<td>373</td>
</tr>
<tr>
<td>A high fever</td>
<td>104</td>
<td>40</td>
<td>313</td>
</tr>
<tr>
<td>Normal body temperature</td>
<td>98.6</td>
<td>37.0</td>
<td>310.</td>
</tr>
<tr>
<td>Room temperature</td>
<td>70</td>
<td>21</td>
<td>294</td>
</tr>
<tr>
<td>Water freezes</td>
<td>32</td>
<td>0</td>
<td>273</td>
</tr>
<tr>
<td>A northern winter</td>
<td>-66</td>
<td>-54</td>
<td>219</td>
</tr>
<tr>
<td>Nitrogen liquefies</td>
<td>-346</td>
<td>-210</td>
<td>63</td>
</tr>
<tr>
<td>Absolute zero</td>
<td>-459</td>
<td>-273</td>
<td>0</td>
</tr>
</tbody>
</table>
SAMPLE PROBLEM 3.3 Converting from Degrees Celsius to Degrees Fahrenheit

The typical temperature in a room is set at 21 °C. What is that temperature in degrees Fahrenheit?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>21 °C</td>
<td>T in degrees Fahrenheit</td>
<td>temperature equation</td>
</tr>
</tbody>
</table>

STEP 2 Write a temperature equation.

\[ T_F = 1.8(T_C) + 32 \]

STEP 3 Substitute in the known values and calculate the new temperature.

\[
T_F = 1.8(21) + 32 \\
= 38 + 32 \\
= 70 °F
\]

Answer to the ones place

In the equation, the values of 1.8 and 32 are exact numbers, which do not affect the number of SFs used in the answer.

STUDY CHECK 3.3

In the process of making ice cream, rock salt is added to crushed ice to chill the ice cream mixture. If the temperature drops to −11 °C, what is it in degrees Fahrenheit?

ANSWER

12 °F

SAMPLE PROBLEM 3.4 Converting from Degrees Fahrenheit to Degrees Celsius

In a type of cancer treatment called *thermotherapy*, temperatures as high as 113 °F are used to destroy cancer cells or make them more sensitive to radiation. What is that temperature in degrees Celsius?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>113 °F</td>
<td>T in degrees Celsius</td>
<td>temperature equation</td>
</tr>
</tbody>
</table>

STEP 2 Write a temperature equation.

\[
T_C = \frac{T_F - 32}{1.8}
\]
**SAMPLE PROBLEM 3.5 Converting from Degrees Celsius to Kelvins**

A dermatologist uses cryogenic liquid nitrogen at −196 °C to remove skin lesions and some skin cancers. What is the temperature, in kelvins, of the liquid nitrogen?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th><strong>ANALYZE THE PROBLEM</strong></th>
<th><strong>Given</strong></th>
<th><strong>Need</strong></th>
<th><strong>Connect</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>−196 °C</td>
<td>T in kelvins</td>
<td>temperature equation</td>
</tr>
</tbody>
</table>

**STEP 2** Write a temperature equation.

\[ T_K = T_C + 273 \]

**STEP 3** Substitute in the known values and calculate the new temperature.

\[ T_K = -196 + 273 \]

\[ = 77 \text{ K} \quad \text{Answer to the ones place} \]

**STUDY CHECK 3.5**

On the planet Mercury, the average night temperature is 13 K and the average day temperature is 683 K. What are these temperatures in degrees Celsius?

**ANSWER**

night −260 °C; day 410 °C

---

**CHEMISTRY LINK TO HEALTH**

**Variation in Body Temperature**

Normal body temperature is considered to be 37.0 °C, although it varies throughout the day and from person to person. Oral temperatures of 36.1 °C are common in the morning and climb to a high of 37.2 °C between 6 P.M. and 10 P.M. Temperatures above 37.2 °C for a person at rest are usually an indication of illness. Individuals who are involved in prolonged exercise may also experience elevated temperatures. Body temperatures of marathon runners can range from 39 °C to 41 °C as heat production during exercise exceeds the body’s ability to lose heat.

Changes of more than 3.5 °C from the normal body temperature begin to interfere with bodily functions. Body temperatures above 41 °C, hyperthermia, can lead to convulsions, particularly in children, which may cause permanent brain damage. Heatstroke occurs above 41.1 °C. Sweat production stops, and the skin becomes hot
3.4 Energy

**LEARNING GOAL** Identify energy as potential or kinetic; convert between units of energy.

Almost everything you do involves energy. When you are running, walking, dancing, or thinking, you are using energy to do work, any activity that requires energy. In fact, energy is defined as the ability to do work. Suppose you are climbing a steep hill and you become too tired to go on. At that moment, you do not have the energy to do any more work. Now suppose you sit down and have lunch. In a while, you will have obtained some energy from the food, and you will be able to do more work and complete the climb.

**Kinetic and Potential Energy**

Energy can be classified as kinetic energy or potential energy. **Kinetic energy** is the energy of motion. Any object that is moving has kinetic energy. **Potential energy** is determined by the position of an object or by the chemical composition of a substance. A boulder...
resting on top of a mountain has potential energy because of its location. If the boulder rolls down the mountain, the potential energy becomes kinetic energy. Water stored in a reservoir has potential energy. When the water goes over the dam and falls to the stream below, its potential energy is converted to kinetic energy. Foods and fossil fuels have potential energy in their molecules. When you digest food or burn gasoline in your car, potential energy is converted to kinetic energy to do work.

**Heat and Energy**

Heat is the energy associated with the motion of particles. An ice cube feels cold because heat flows from your hand into the ice cube. The faster the particles move, the greater the heat or thermal energy of the substance. In the ice cube, the particles are moving very slowly. As heat is added, the motion of the particles in the ice cube increases. Eventually, the particles have enough energy to make the ice cube melt as it changes from a solid to a liquid.

**Units of Energy**

The SI unit of energy and work is the joule (J) (pronounced “jewel”). The joule is a small amount of energy, so scientists often use the kilojoule (kJ), 1000 joules. To heat water for one cup of tea, you need about 75 000 J or 75 kJ of heat. TABLE 3.6 shows a comparison of energy in joules for several energy sources or uses.

You may be more familiar with the unit calorie (cal), from the Latin caloric, meaning “heat.” The calorie was originally defined as the amount of energy (heat) needed to raise the temperature of 1 g of water by 1 °C. Now, one calorie is defined as exactly 4.184 J. This equality can be written as two conversion factors:

\[
1 \text{ cal} = 4.184 \text{ J (exact)}
\]

\[
1 \text{ cal} = \frac{4.184 \text{ J}}{1 \text{ cal}}
\]

\[
\frac{4.184 \text{ J}}{1 \text{ cal}}
\]

\[
\frac{1 \text{ kcal}}{1000 \text{ cal}}
\]

One kilocalorie (kcal) is equal to 1000 calories, and one kilojoule (kJ) is equal to 1000 joules. The equalities and conversion factors follow:

\[
1 \text{ kcal} = 1000 \text{ cal}
\]

\[
\frac{1000 \text{ cal}}{1 \text{ kcal}}
\]

\[
\frac{1 \text{ kcal}}{1000 \text{ cal}}
\]

\[
1 \text{ kJ} = 1000 \text{ J}
\]

\[
\frac{1000 \text{ J}}{1 \text{ kJ}}
\]

\[
\frac{1 \text{ kJ}}{1000 \text{ J}}
\]

**TABLE 3.6 A Comparison of Energy for Various Resources and Uses**

<table>
<thead>
<tr>
<th>Energy in Joules</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>(10^9)</td>
<td>Energy use per person in 1 yr in the United States (10^11)</td>
</tr>
<tr>
<td>(10^8)</td>
<td>Energy from 1 gal of gasoline (10^8)</td>
</tr>
<tr>
<td>(10^7)</td>
<td>Energy from one serving of pasta, a doughnut, or needed to bicycle for 1 h (10^6)</td>
</tr>
<tr>
<td>(10^6)</td>
<td>Energy used to sleep for 1 h (10^5)</td>
</tr>
<tr>
<td>(10^5)</td>
<td>Solar energy reaching the Earth in 1 s (10^17)</td>
</tr>
<tr>
<td>(10^4)</td>
<td>Energy consumption for 1 yr in the United States (10^20)</td>
</tr>
<tr>
<td>(10^3)</td>
<td>World reserves of fossil fuel (10^23)</td>
</tr>
<tr>
<td>(10^2)</td>
<td>Energy radiated by the Sun in 1 s (10^26)</td>
</tr>
<tr>
<td>(10^0)</td>
<td>-</td>
</tr>
</tbody>
</table>

**CORE CHEMISTRY SKILL**

Using Energy Units
SAMPLE PROBLEM 3.6 Energy Units
A defibrillator gives a high-energy-shock output of 360 J. What is this quantity of energy in calories?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>360 J</td>
<td>calories</td>
<td>energy factor</td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to convert the given unit to the needed unit.

joules $\frac{\text{energy factor}}{\text{calories}}$

STEP 3 State the equalities and conversion factors.

$1 \text{ cal} = 4.184 \text{ J}$

$\frac{4.184 \text{ J}}{1 \text{ cal}}$ and $\frac{1 \text{ cal}}{4.184 \text{ J}}$

STEP 4 Set up the problem to calculate the needed quantity.

$360 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 86 \text{ cal}$

Two SFs Exact Two SFs

STUDY CHECK 3.6
When 1.0 g of glucose is metabolized in the body, it produces 3.9 kcal. How many kilojoules are produced?

ANSWER
16 kJ

QUESTIONS AND PROBLEMS

3.4 Energy

LEARNING GOAL Identify energy as potential or kinetic; convert between units of energy.

3.21 Discuss the changes in the potential and kinetic energy of a roller-coaster ride as the roller-coaster car climbs to the top and goes down the other side.

3.22 Discuss the changes in the potential and kinetic energy of a ski jumper taking the elevator to the top of the jump and going down the ramp.

3.23 Indicate whether each of the following statements describes potential or kinetic energy:
   a. water at the top of a waterfall
   b. kicking a ball
   c. the energy in a lump of coal
   d. a skier at the top of a hill

3.24 Indicate whether each of the following statements describes potential or kinetic energy:
   a. the energy in your food
   b. a tightly wound spring
   c. an earthquake
   d. a car speeding down the freeway

3.25 Convert each of the following energy units:
   a. 3500 cal to kcal
   b. 415 J to cal
   c. 28 cal to J
   d. 4.5 kJ to cal

3.26 Convert each of the following energy units:
   a. 8.1 kcal to cal
   b. 325 J to kJ
   c. 2550 cal to kJ
   d. 2.50 kcal to J
Applications

3.27 The energy needed to keep a 75-watt light bulb burning for 1.0 h is 270 kJ. Calculate the energy required to keep the light bulb burning for 3.0 h in each of the following energy units:
   a. joules
   b. kilocalories

3.28 A person uses 750 kcal on a long walk. Calculate the energy used for the walk in each of the following energy units:
   a. joules
   b. kilojoules

CHEMISTRY LINK TO THE ENVIRONMENT
Carbon Dioxide and Climate Change

The Earth’s climate is a product of interactions between sunlight, the atmosphere, and the oceans. The Sun provides us with energy in the form of solar radiation. Some of this radiation is reflected back into space. The rest is absorbed by the clouds, atmospheric gases including carbon dioxide (CO₂), and the Earth’s surface. For millions of years, concentrations of carbon dioxide have fluctuated. However in the last 100 years, the amount of CO₂ gas in our atmosphere has increased significantly. From the years 1000 to 1800, the atmospheric carbon dioxide level averaged 280 ppm. The concentration of ppm indicates the parts per million by volume, which for gases is the same as mL of CO₂ per kL of air. But since the beginning of the Industrial Revolution in 1800 up until 2005, the level of atmospheric carbon dioxide has risen from about 280 ppm to about 394 ppm in 2013, a 40% increase.

As the atmospheric CO₂ level increases, more solar radiation is trapped by the atmospheric gases, which raises the temperature at the surface of the Earth. Some scientists have estimated that if the carbon dioxide level doubles from its level before the Industrial Revolution, the average global temperature could increase by 2.0 to 4.4°C. Although this seems to be a small temperature change, it could have dramatic impact worldwide. Even now, glaciers and snow cover in much of the world have diminished. Ice sheets in Antarctica and Greenland are melting faster and breaking apart. Although no one knows for sure how rapidly the ice in the polar regions is melting, this accelerating change will contribute to a rise in sea level. In the twentieth century, the sea level rose by 15 to 23 cm. Some scientists predict the sea level will rise by 1 m in this century. Such an increase will have a major impact on coastal areas.

Until recently, the carbon dioxide level was maintained as algae in the oceans and the trees in the forests utilized the carbon dioxide. However, the ability of these and other forms of plant life to absorb carbon dioxide is not keeping up with the increase in carbon dioxide. Most scientists agree that the primary source of the increase of carbon dioxide is the burning of fossil fuels such as gasoline, coal, and natural gas. The cutting and burning of trees in the rainforests (deforestation) also reduces the amount of carbon dioxide removed from the atmosphere.

Worldwide efforts are being made to reduce the carbon dioxide produced by burning fossil fuels that heat our homes, run our cars, and provide energy for industries. Scientists are exploring ways to provide alternative energy sources and to reduce the effects of deforestation. Meanwhile, we can reduce energy use in our homes by using appliances that are more energy efficient such as replacing incandescent light bulbs with fluorescent lights. Such an effort worldwide will reduce the possible impact of climate change and at the same time save our fuel resources.
3.5 Specific Heat

**LEARNING GOAL** Calculate the specific heat for a substance. Use specific heat to calculate heat loss or gain.

Every substance has its own characteristic ability to absorb heat. When you bake a potato, you place it in a hot oven. If you are cooking pasta, you add the pasta to boiling water. You already know that adding heat to water increases its temperature until it boils. Certain substances must absorb more heat than others to reach a certain temperature.

The energy requirements for different substances are described in terms of a physical property called specific heat. The specific heat (SH) for a substance is defined as the amount of heat (q) in calories (or joules) needed to change the temperature of exactly 1 g of a substance by exactly 1 °C. To calculate the specific heat for a substance, we measure the heat in calories or joules, the mass in grams, and the temperature change written as ΔT.

\[
SH = \frac{q}{m \times \Delta T} = \frac{\text{cal (or J)}}{\text{g °C}}
\]

The specific heat for water is written using our definition of the calorie and joule.

\[
SH \text{ for H}_2\text{O(l)} = \frac{1.00 \text{ cal}}{\text{g °C}} = \frac{4.184 \text{ J}}{\text{g °C}}
\]

If we look at **TABLE 3.7**, we see that 1 g of water requires 1.00 cal or 4.184 J to increase its temperature by 1 °C. Water has a large specific heat that is about five times the specific heat of aluminum. Aluminum has a specific heat that is about twice that of copper. However, adding the same amount of heat (1.00 cal or 4.184 J) will raise the temperature of 1 g of aluminum by about 5 °C and 1 g of copper by about 10 °C. The low specific heats of aluminum and copper mean they transfer heat efficiently, which makes them useful in cookware.

The high specific heat of water has a major impact on the temperatures in a coastal city compared to an inland city. A large mass of water near a coastal city can absorb or release five times the energy absorbed or released by the same mass of rock near an inland city. This means that in the summer a body of water absorbs large quantities of heat, which cools a coastal city, and then in the winter that same body of water releases large quantities of heat, which provides warmer temperatures. A similar effect happens with our bodies, which contain 70% water by mass. Water in the body absorbs or releases large quantities of heat to maintain an almost constant body temperature.

**SAMPLE PROBLEM 3.7 Calculating Specific Heat**

What is the specific heat, in J/g °C, of lead if 57.0 J raises the temperature of 35.6 g of lead by 12.5 °C?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>( q = 57.0 \text{ J} )</td>
<td>( m = 35.6 \text{ g of lead} )</td>
<td>specific heat of lead (J/g °C)</td>
</tr>
<tr>
<td>( \Delta T = 12.5 \text{ °C} )</td>
<td></td>
<td>specific heat equation</td>
</tr>
</tbody>
</table>
**Guide to Calculating Specific Heat**

**STEP 1**
State the given and needed quantities.

**STEP 2**
Write the relationship for specific heat.

**STEP 3**
Set up the problem to calculate the specific heat.

---

**CORE CHEMISTRY SKILL**
Calculating Specific Heat

**CORE CHEMISTRY SKILL**
Using the Heat Equation

---

**STEP 2** Write the relationship for specific heat. The specific heat ($SH$) is calculated by dividing the heat ($q$) by the mass ($m$) and by the temperature change ($\Delta T$).

$$SH = \frac{\text{heat}}{\text{mass} \cdot \Delta T} = \frac{q}{m \cdot \Delta T}$$

**STEP 3** Set up the problem to calculate the specific heat.

Three SFs

$$SH = \frac{57.0 \text{ J}}{35.6 \text{ g} \cdot 12.5 ^\circ \text{C}} = \frac{0.128 \text{ J}}{\text{g} \cdot ^\circ \text{C}}$$

**STUDY CHECK 3.7**
What is the specific heat, in $\text{J/g} \cdot ^\circ \text{C}$, of sodium if 92.3 J is needed to raise the temperature of 3.00 g of sodium by 25.0 °C?

**ANSWER**
$1.23 \text{ J/g} \cdot ^\circ \text{C}$

---

**Heat Equation**

When we know the specific heat of a substance, we can calculate the heat lost or gained by measuring the mass of the substance and the initial and final temperatures. We can substitute these measurements into the specific heat equation that is rearranged to solve for heat, which we call the heat equation.

$$SH = \frac{q}{m \times \Delta T}$$

$$m \times \Delta T \times SH = \frac{q}{\rho \times \Delta T} \times \rho \times \Delta T$$

Now we can write the heat equation as

$$q = m \times \Delta T \times SH$$

The heat lost or gained, in calories or joules, is obtained when the units of grams and °C in the numerator cancel grams and °C in the denominator of specific heat in the heat equation.

$$\text{cal} = \text{g} \times ^\circ \text{C} \times \frac{\text{cal}}{\text{g} \cdot ^\circ \text{C}}$$

$$\text{J} = \text{g} \times ^\circ \text{C} \times \frac{\text{J}}{\text{g} \cdot ^\circ \text{C}}$$

---

**SAMPLE PROBLEM 3.8 Calculating Heat Loss**

During surgery or when a patient has suffered a cardiac arrest or stroke, lowering the body temperature will reduce the amount of oxygen needed by the body. Some methods used to lower body temperature include cooled saline solution, cool water blankets, or cooling caps worn on the head. How many kilojoules are lost when the body temperature of a surgery patient with a blood volume of 5500 mL is cooled from 38.5 °C to 33.2 °C? (Assume that the specific heat and density of blood is the same as for water.)
**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>5500 mL of blood</td>
<td></td>
<td>kilojoules</td>
<td>specific heat of water</td>
</tr>
<tr>
<td>( = 5500 \text{ g of blood,} )</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>( = 38.5 \text{ °C to 33.2 °C} )</td>
<td>removed</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Calculate the temperature change (\( \Delta T \)).

\[
\Delta T = 38.5 \text{ °C} - 33.2 \text{ °C} = 5.3 \text{ °C}
\]

**STEP 3** Write the heat equation and needed conversion factors.

\[
q = m \times \Delta T \times SH
\]

\[
SH_{water} = \frac{4.184 \text{ J}}{\text{g °C}}
\]

\[
1 \text{ kJ} = 1000 \text{ J}
\]

**STEP 4** Substitute in the given values and calculate the heat, making sure units cancel.

\[
q = 5500 \times 5.3 \times 4.184 \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 120 \text{ kJ}
\]

**STUDY CHECK 3.8**

Some cooking pans have a layer of copper on the bottom. How many kilojoules are needed to raise the temperature of 125 g of copper from 22 °C to 325 °C (see Table 3.7)?

**ANSWER**

14.6 kJ

Another use of the heat equation is to calculate the mass, in grams, of a substance by rearranging the heat equation to solve for mass (\( m \)) as shown in Sample Problem 3.9.

**SAMPLE PROBLEM 3.9 Using the Heat Equation**

When 655 J is added to a sample of ethanol, its temperature rises from 18.2 °C to 32.8 °C. What is the mass, in grams, of the ethanol sample (see Table 3.7)?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>( q = 655 \text{ J,} )</td>
<td>grams of ethanol</td>
<td></td>
<td>heat equation</td>
</tr>
<tr>
<td>heated from 18.2 °C to 32.8 °C</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Calculate the temperature change (\( \Delta T \)).

\[
\Delta T = 32.8 \text{ °C} - 18.2 \text{ °C} = 14.6 \text{ °C}
\]
**Step 3** Write the heat equation.

\[ q = m \times \Delta T \times SH \]

When the heat equation is rearranged for mass \( m \), the heat is divided by the temperature change and the specific heat.

\[ m = \frac{q}{\Delta T \times SH} \]

**Step 4** Substitute in the given values and solve, making sure units cancel.

\[ \frac{655 \text{ J}}{14.6 \degree \text{C}} \times \frac{2.46 \text{ J}}{\text{g} \degree \text{C}} = 18.2 \text{ g} \]

**Study Check 3.9**
When 8.81 kJ is absorbed by a piece of iron, its temperature rises from 15 °C to 122 °C. What is the mass, in grams, of the piece of iron (see Table 3.7)?

**Answer**
182 g

**Questions and Problems**

**3.5 Specific Heat**

**Learning Goal** Calculate the specific heat for a substance. Use specific heat to calculate heat loss or gain, temperature change, or mass of a sample.

**3.29** If the same amount of heat is supplied to samples of 10.0 g each of aluminum, iron, and copper all at 15.0 °C, which sample would reach the highest temperature (see Table 3.7)?

**3.30** Substances A and B are the same mass and at the same initial temperature. When the same amount of heat is added to each, the final temperature of A is 75 °C and B is 35 °C. What does this tell you about the specific heats of A and B?

**3.31** Calculate the specific heat \((J/\text{g} °\text{C})\) for each of the following:

- a. a 13.5-g sample of zinc (Zn) heated from 24.2 °C to 83.6 °C that absorbs 312 J of heat
- b. a 48.2-g sample of a metal that absorbs 345 J with a temperature increase from 35.0 °C to 57.9 °C

**3.32** Calculate the specific heat \((J/\text{g} °\text{C})\) for each of the following:

- a. an 18.5-g sample of tin (Sn) that absorbs 183 J of heat when its temperature increases from 35.0 °C to 78.6 °C
- b. a 22.5-g sample of a metal that absorbs 645 J when its temperature increases from 36.2 °C to 92.0 °C

**3.33** Use the heat equation to calculate the energy, in joules, for each of the following (see Table 3.7):

- a. required to heat 25.0 g of water, \(H_2O\), from 5.5 °C to 25.7 °C
- b. required to heat 38.0 g of copper (Cu) from 122 °C to 246 °C

**3.34** Use the heat equation to calculate the energy, in joules, for each of the following (see Table 3.7):

- a. required to heat 5.25 g of water, \(H_2O\), from 5.5 °C to 64.8 °C
- b. required to heat 10.0 g of silver (Ag) from 112 °C to 275 °C

**3.35** Calculate the mass, in grams, for each of the following using Table 3.7:

- a. a sample of gold (Au) that absorbs 225 J to increase its temperature from 15.0 °C to 47.0 °C
- b. a sample of iron (Fe) that loses 8.40 kJ when its temperature decreases from 168.0 °C to 82.0 °C

**3.36** Calculate the mass, in grams, for each of the following using Table 3.7:

- a. a sample of water, \(H_2O\), that absorbs 8250 J when its temperature increases from 18.4 °C to 92.6 °C
- b. a sample of silver (Ag) that loses 3.22 kJ when its temperature decreases from 145 °C to 24 °C

**3.37** Calculate the change in temperature \((\Delta T)\) for each of the following using Table 3.7:

- a. 20.0 g of iron (Fe) that absorbs 1580 J
- b. 150.0 g of water, \(H_2O\), that absorbs 7.10 kJ

**3.38** Calculate the change in temperature \((\Delta T)\) for each of the following using Table 3.7:

- a. 115 g of copper (Cu) that loses 2.45 kJ
- b. 22.0 g of silver (Ag) that loses 625 J
3.6 Energy and Nutrition

LEARNING GOAL Use the energy values to calculate the kilocalories (kcal) or kilojoules (kJ) for a food.

The food we eat provides energy to do work in the body, which includes the growth and repair of cells. Carbohydrates are the primary fuel for the body, but if the carbohydrate reserves are exhausted, fats and then proteins are used for energy.

For many years in the field of nutrition, the energy from food was measured as Calories or kilocalories. The nutritional unit Calorie, Cal (with an uppercase C), is the same as 1000 cal, or 1 kcal. The international unit, kilojoule (kJ), is becoming more prevalent. For example, a baked potato has an energy value of 100 Calories, which is 100 kcal or 440 kJ. A typical diet that provides 2100 Cal (kcal) is the same as an 8800 kJ diet.

\[ 1 \text{ Cal} = 1 \text{ kcal} = 1000 \text{ cal} \]
\[ 1 \text{ Cal} = 4.184 \text{ kJ} = 4184 \text{ J} \]

In the nutrition laboratory, foods are burned in a calorimeter to determine their energy value (kcal/g or kJ/g)(see FIGURE 3.6). A sample of food is placed in a steel container called a calorimeter filled with oxygen. A measured amount of water is added to fill the area surrounding the combustion chamber. The food sample is ignited, releasing heat that increases the temperature of the water. From the known mass of the food and water as well as the measured temperature increase, the energy value for the food is calculated. We will assume that the energy absorbed by the calorimeter is negligible.

![Calorimeter Diagram](image)

**FIGURE 3.6** Heat released from burning a food sample in a calorimeter is used to determine the energy value for the food.

What happens to the temperature of water in a calorimeter during the combustion of a food sample?

**Energy Values for Foods**

The energy values for food are the kilocalories or kilojoules obtained from burning 1 g of carbohydrate, fat, or protein, which are listed in TABLE 3.8. Using these energy values, we can calculate the total energy for a food if the mass of each food type is known.

\[ \text{kilocalories} = g \times \frac{\text{kcal}}{g} \quad \text{kilojoules} = g \times \frac{\text{kJ}}{g} \]

On packaged food, the energy content is listed on the Nutrition Facts label, usually in terms of the number of Calories or kilojoules for one serving. The general composition and energy content for some foods are given in TABLE 3.9. The total energy, in kilocalories, for each food type was calculated using energy values in kilocalories. Total energy, in kilojoules, was calculated using energy values in kilojoules. The energy for each food type was rounded off to the tens place.

<table>
<thead>
<tr>
<th>Food Type</th>
<th>kcal/g</th>
<th>kJ/g</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbohydrate</td>
<td>4</td>
<td>17</td>
</tr>
<tr>
<td>Fat</td>
<td>9</td>
<td>38</td>
</tr>
<tr>
<td>Protein</td>
<td>4</td>
<td>17</td>
</tr>
</tbody>
</table>

What type of food provides the most energy per gram for the human body?
SAMPLE PROBLEM 3.10 Energy Content for a Food Using Nutrition Facts

The Nutrition Facts label for crackers states that 1 serving contains 19 g of carbohydrate, 4 g of fat, and 2 g of protein. What is the energy, in kilocalories, from each food type and the total kilocalories for one serving of crackers? Round off the kilocalories for each food type to the tens place.

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

Given:
- 19 g of carbohydrate
- 4 g of fat
- 2 g of protein

Need:
- Total number of kilocalories

Connect:
- Energy values

ANALYZE THE PROBLEM

STEP 2 Use the energy value for each food type and calculate the kcal rounded off to the tens place. Using the energy values for carbohydrate, fat, and protein (see Table 3.8), we can calculate the energy for each type of food.

<table>
<thead>
<tr>
<th>Food Type</th>
<th>Mass (g)</th>
<th>Energy Value (kcal/1 g)</th>
<th>Energy (kcal)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbohydrate</td>
<td>19</td>
<td>4</td>
<td>80</td>
</tr>
<tr>
<td>Fat</td>
<td>4</td>
<td>9</td>
<td>40</td>
</tr>
<tr>
<td>Protein</td>
<td>2</td>
<td>4</td>
<td>10</td>
</tr>
</tbody>
</table>

STEP 3 Add the energy for each food type to give the total energy from the food.

Total energy = 80 kcal + 40 kcal + 10 kcal = 130 kcal
STUDY CHECK 3.10

a. Using the nutrition values for one serving of crackers in Sample Problem 3.10, calculate the energy, in kilojoules, for each food type. Round off the kilojoules for each food type to the tens place.
b. What is the total energy, in kilojoules, for one serving of crackers?

ANSWER

a. carbohydrate, 630 kJ; fat, 300 kJ; protein, 50 kJ
b. 980 kJ

CHEMISTRY LINK TO HEALTH

Losing and Gaining Weight

The number of kilocalories or kilojoules needed in the daily diet of an adult depends on gender, age, and level of physical activity. Some typical levels of energy needs are given in TABLE 3.10.

<table>
<thead>
<tr>
<th>TABLE 3.10 Typical Energy Requirements for Adults</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gender</td>
</tr>
<tr>
<td>Female</td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td>Male</td>
</tr>
<tr>
<td></td>
</tr>
</tbody>
</table>

A person gains weight when food intake exceeds energy output. The amount of food a person eats is regulated by the hunger center in the hypothalamus, which is located in the brain. Food intake is normally proportional to the nutrient stores in the body. If these nutrient stores are low, you feel hungry; if they are high, you do not feel like eating.

A person loses weight when food intake is less than energy output. Many diet products contain cellulose, which has no nutritive value but provides bulk and makes you feel full. Some diet drugs depress the hunger center and must be used with caution, because they excite the nervous system and can elevate blood pressure. Because muscular exercise is an important way to expend energy, an increase in daily exercise aids weight loss. TABLE 3.11 lists some activities and the amount of energy they require.

<table>
<thead>
<tr>
<th>TABLE 3.11 Energy Expended by a 70.0-kg (154-lb) Adult</th>
</tr>
</thead>
<tbody>
<tr>
<td>Activity</td>
</tr>
<tr>
<td>Sleeping</td>
</tr>
<tr>
<td>Sitting</td>
</tr>
<tr>
<td>Walking</td>
</tr>
<tr>
<td>Swimming</td>
</tr>
<tr>
<td>Running</td>
</tr>
</tbody>
</table>

One hour of swimming uses 2100 kJ of energy.

QUESTIONS AND PROBLEMS

3.6 Energy and Nutrition

LEARNING GOAL Use the energy values to calculate the kilocalories (kcal) or kilojoules (kJ) for a food.

3.39 Calculate the kilocalories for each of the following:
a. one stalk of celery that produces 125 kJ when burned in a calorimeter
b. a waffle that produces 870. kJ when burned in a calorimeter

3.40 Calculate the kilocalories for each of the following:
a. one cup of popcorn that produces 131 kJ when burned in a calorimeter
b. a sample of butter that produces 23.4 kJ when burned in a calorimeter
3.41 Using the energy values for foods (see Table 3.8), determine each of the following (round off the answer for each food type to the tens place):

a. the total kilojoules for one cup of orange juice that contains 26 g of carbohydrate, no fat, and 2 g of protein
b. the grams of carbohydrate in one apple if the apple has no fat and no protein and provides 72 kcal of energy
c. the kilocalories in one tablespoon of vegetable oil, which contains 14 g of fat and no carbohydrate or protein
d. the grams of fat in one avocado that has 405 kcal, 13 g of carbohydrate, and 5 g of protein

3.42 Using the energy values for foods (see Table 3.8), determine each of the following (round off the answer for each food type to the tens place):

a. the total kilojoules in two tablespoons of crunchy peanut butter that contains 6 g of carbohydrate, 16 g of fat, and 7 g of protein
b. the grams of protein in one cup of soup that has 110 kcal with 9 g of carbohydrate and 6 g of fat
c. the kilocalories in one can of cola if it has 40 g of carbohydrate and no fat or protein
d. the total kilocalories for a diet that consists of 68 g of carbohydrate, 9 g of fat, and 150 g of protein

3.43 One cup of clam chowder contains 16 g of carbohydrate, 12 g of fat, and 9 g of protein. How much energy, in kilocalories and kilojoules, is in the clam chowder? (Round off the answer for each food type to the tens place.)

3.44 A high-protein diet contains 70.0 g of carbohydrate, 5.0 g of fat, and 150 g of protein. How much energy, in kilocalories and kilojoules, does this diet provide? (Round off the answer for each food type to the tens place.)

Applications

3.45 A patient receives 3.2 L of glucose solution intravenously (IV). If 100 mL of the solution contains 5.0 g of glucose (carbohydrate), how many kilocalories did the patient obtain from the glucose solution?

3.46 For lunch, a patient consumed 3 oz of skinless chicken, 3 oz of broccoli, 1 medium apple, and 1 cup of nonfat milk (see Table 3.9). How many kilocalories did the patient obtain from the lunch?

Follow Up

A DIET AND EXERCISE PROGRAM FOR CHARLES

It has been two weeks since Charles met with Daniel, a dietitian, who provided Charles with a menu for weight loss. Charles and his mother are going back to see Daniel again with a chart of the food Charles has eaten. The following is what Charles ate in one day:

Breakfast
1 banana, 1 cup of nonfat milk, 1 egg

Lunch
1 cup of carrots, 3 oz of ground beef, 1 apple, 1 cup of nonfat milk

Dinner
6 oz of skinless chicken, 1 baked potato, 3 oz of broccoli, 1 cup of nonfat milk

Applications

3.47 Using energy values from Table 3.9, determine each of the following:

a. the total kilocalories for each meal
b. the total kilocalories for one day
c. If Charles consumes 1800 kcal per day, he will maintain his weight. Would he lose weight on his new diet?
d. If expending 3500 kcal is equal to a loss of 1.0 lb, how many days will it take Charles to lose 5.0 lb?

3.48 a. During one week, Charles swam for a total of 2.5 h and walked for a total of 8.0 h. If Charles expends 340 kcal/h swimming and 160 kcal/h walking, how many total kilocalories did he expend for one week?
b. For the amount of exercise that Charles did for one week in part a, if expending 3500 kcal is equal to a loss of 1.0 lb, how many pounds did he lose?
c. How many hours would Charles have to walk to lose 1.0 lb?
d. How many hours would Charles have to swim to lose 1.0 lb?
3.1 Classification of Matter

**LEARNING GOAL** Classify examples of matter as pure substances or mixtures.

- Matter is anything that has mass and occupies space.
- Matter is classified as pure substances or mixtures.
- Pure substances, which are elements or compounds, have fixed compositions, and mixtures have variable compositions.
- The substances in mixtures can be separated using physical methods.

3.2 States and Properties of Matter

**LEARNING GOAL** Identify the states and the physical and chemical properties of matter.

- The three states of matter are solid, liquid, and gas.
- A physical property is a characteristic of a substance that can be observed or measured without affecting the identity of the substance.
- A physical change occurs when physical properties change, but not the composition of the substance.
- A chemical property indicates the ability of a substance to change into another substance.
- A chemical change occurs when one or more substances react to form a substance with new physical and chemical properties.

3.3 Temperature

**LEARNING GOAL** Given a temperature, calculate a corresponding temperature on another scale.

- In science, temperature is measured in degrees Celsius (°C) or kelvins (K).
- On the Celsius scale, there are 100 units between the freezing point (0 °C) and the boiling point (100 °C) of water.
- On the Fahrenheit scale, there are 180 units between the freezing point (32 °F) and the boiling point (212 °F) of water. A Fahrenheit temperature is related to its Celsius temperature by the equation \( T_F = \frac{9}{5}T_C + 32 \).
- The SI unit, kelvin, is related to the Celsius temperature by the equation \( T_K = T_C + 273 \).

3.4 Energy

**LEARNING GOAL** Identify energy as potential or kinetic; convert between units of energy.

- Energy is the ability to do work.
- Potential energy is stored energy; kinetic energy is the energy of motion.
- Common units of energy are the calorie (cal), kilocalorie (kcal), joule (J), and kilojoule (kJ).
- One calorie is equal to 4.184 J.
3.5 Specific Heat

**LEARNING GOAL** Calculate the specific heat for a substance. Use specific heat to calculate heat loss or gain.

- Specific heat is the amount of energy required to raise the temperature of exactly 1 g of a substance by exactly 1 °C.
- The heat lost or gained by a substance is determined by multiplying its mass, the temperature change, and its specific heat.

### KEY TERMS

- **calorie (cal)** The amount of heat energy that raises the temperature of exactly 1 g of water by exactly 1 °C.
- **change of state** The transformation of one state of matter to another; for example, solid to liquid, liquid to solid, liquid to gas.
- **chemical change** A change during which the original substance is converted into a new substance that has a different composition and new physical and chemical properties.
- **chemical properties** The properties that indicate the ability of a substance to change into a new substance.
- **compound** A pure substance consisting of two or more elements, with a definite composition, that can be broken down into simpler substances only by chemical methods.
- **element** A pure substance containing only one type of matter, which cannot be broken down by chemical methods.
- **energy** The ability to do work.
- **energy value** The kilocalories (or kilojoules) obtained per gram of the food types: carbohydrate, fat, and protein.
- **gas** A state of matter that does not have a definite shape or volume.
- **heat** The energy associated with the motion of particles in a substance.
- **heat equation** A relationship that calculates heat \( q \) given the mass, specific heat, and temperature change for a substance.
- **joule (J)** The SI unit of heat energy; 4.184 J = 1 cal.
- **kinetic energy** The energy of moving particles.
- **liquid** A state of matter that takes the shape of its container but has a definite volume.
- **matter** The material that makes up a substance and has mass and occupies space.
- **mixture** The physical combination of two or more substances that does not change the identities of the mixed substances.
- **physical change** A change in which the physical properties of a substance change but its identity stays the same.
- **physical properties** The properties that can be observed or measured without affecting the identity of a substance.
- **potential energy** A type of energy related to position or composition of a substance.
- **pure substance** A type of matter that has a definite composition.
- **solid** A state of matter that has its own shape and volume.
- **specific heat (SH)** A quantity of heat that changes the temperature of exactly 1 g of a substance by exactly 1 °C.
- **states of matter** Three forms of matter: solid, liquid, and gas.

### CORE CHEMISTRY SKILLS

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

**Classifying Matter (3.1)**

- A pure substance is matter that has a fixed or constant composition.
- An element, the simplest type of a pure substance, is composed of only one type of matter, such as silver, iron, or aluminum.
- A compound is also a pure substance, but it consists of two or more elements chemically combined in the same proportion.
- In a **homogeneous mixture**, also called a solution, the composition of the substances in the mixture is uniform.
- In a **heterogeneous mixture**, the components are visible and do not have a uniform composition throughout the sample.

**Example:** Classify each of the following as a pure substance (element or compound) or a mixture (homogeneous or heterogeneous):

- a. iron in a nail
- b. black coffee
- c. carbon dioxide, a greenhouse gas

**Answer:**

- a. Iron is a pure substance, which is an element.
- b. Black coffee contains different substances with uniform composition, which makes it a homogeneous mixture.
- c. The gas carbon dioxide, which is a pure substance that contains two elements chemically combined, is a compound.

### Identifying Physical and Chemical Changes (3.2)

- When matter undergoes a physical change, its state or its appearance changes, but its composition remains the same.
- When a chemical change takes place, the original substance is converted into a new substance, which has different physical and chemical properties.

**Example:** Classify each of the following as a physical or chemical change:

- a. Silver metal is melted and poured into a mold to make a ring.
- b. Methane, in natural gas, burns in air.
- c. Potassium reacts spontaneously with water to form hydrogen gas.

**Answer:**

- a. A change from a solid to a liquid (melting) is a physical change.
- b. When methane burns, it changes to different substances with new properties, which is a chemical change.
- c. The reaction of potassium with water produces new substances (including hydrogen gas), which makes this a chemical change.
Converting between Temperature Scales (3.3)

- The temperature equation \( T_F = 1.8(T_C) + 32 \) is used to convert from Celsius to Fahrenheit and can be rearranged to convert from Fahrenheit to Celsius.
- The temperature equation \( T_K = T_C + 273 \) is used to convert from Celsius to Kelvin and can be rearranged to convert from Kelvin to Celsius.

**Example:** Convert 75.0 °C to degrees Fahrenheit.
**Answer:**
\[
T_F = 1.8(T_C) + 32
\]
\[
T_F = 1.8(75.0) + 32 = 135 + 32 = 167 \text{ °F}
\]

**Example:** Convert 355 K to degrees Celsius.
**Answer:**
\[
T_K = T_C + 273
\]
To solve the equation for \( T_C \), subtract 273 from both sides.
\[
T_K - 273 = T_C + 273 - 273
\]
\[
T_C = 355 - 273 = 82 \text{ °C}
\]

Using Energy Units (3.4)

- Equalities for energy units include 1 cal = 4.184 J, 1 kcal = 1000 cal, and 1 kJ = 1000 J.
- Each equality for energy units can be written as two conversion factors:
\[
\begin{align*}
\text{1 cal} & = \frac{4.184 \text{ J}}{1 \text{ cal}} & \text{1 kcal} & = \frac{1000 \text{ cal}}{1 \text{ kcal}} \\
\text{1 kJ} & = \frac{1000 \text{ J}}{1 \text{ kJ}}
\end{align*}
\]
- The energy unit conversion factors are used to cancel given units of energy and to obtain the needed unit of energy.

**Example:** Convert 45 000 J to kilocalories.
**Answer:** Using the conversion factors above, we start with the given 45 000 J and convert it to kilocalories.
\[
45 000 \text{ J} \times \frac{1 \text{ kcal}}{4.184 \text{ J}} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} \times \frac{1 \text{ kcal}}{1000 \text{ J}} = 11 \text{ kcal}
\]

Calculating Specific Heat (3.5)

- Specific heat \( (SH) \) is the amount of heat \( (q) \) that raises the temperature of 1 g of a substance by 1 °C.
\[
SH = \frac{q}{m \times \Delta T} \quad \text{cal (or J)} \quad \text{g °C}
\]
- To calculate specific heat, the heat lost or gained is divided by the mass of the substance and the change in temperature \( (\Delta T) \).

**Example:** Calculate the specific heat, in J/g °C, for a 4.0-g sample of tin that absorbs 66 J when heated from 125 °C to 197 °C.
**Answer:**
\[
SH = \frac{q}{m \times \Delta T} = \frac{66 \text{ J}}{4.0 \text{ g} \times 72 \text{ °C}} = 0.23 \text{ J/g °C}
\]

Using the Heat Equation (3.5)

- The quantity of heat \( (q) \) absorbed or lost by a substance is calculated using the heat equation.
\[
q = m \times \Delta T \times SH
\]
- Heat, in calories or joules, is obtained when the specific heat of a substance is used.
- To cancel, the unit grams is used for mass, and the unit °C is used for temperature change.

**Example:** How many joules are required to heat 5.25 g of titanium from 85.5 °C to 132.5 °C?
**Answer:**
\[
m = 5.25 \text{ g} \quad \Delta T = 132.5 \text{ °C} - 85.5 \text{ °C} = 47.0 \text{ °C}
\]
\[
SH \text{ for titanium} = 0.523 \text{ J/g °C}
\]
The known values are substituted into the heat equation making sure units cancel.
\[
q = m \times \Delta T \times SH = 5.25 \text{ g} \times 47.0 \text{ °C} \times \frac{0.523 \text{ J}}{\text{g °C}} = 129 \text{ J}
\]

**Understanding the Concepts**

The chapter sections to review are shown in parentheses at the end of each question.

**3.49** Identify each of the following as an element, a compound, or a mixture. Explain your choice. (3.1)

a. ![Element](image1.png)

b. ![Compound](image2.png)

c. ![Mixture](image3.png)

**3.50** Identify each of the following as a homogeneous or heterogeneous mixture. Explain your choice. (3.1)

a. ![Homogeneous](image4.png)

b. ![Heterogeneous](image5.png)
3.51 Classify each of the following as a homogeneous or heterogeneous mixture: (3.1)
   a. pulpy orange juice
   b. cereal in milk
   c. copper sulphate in water

3.52 Classify each of the following as a homogeneous or heterogeneous mixture: (3.1)
   a. ketchup
   b. hard-boiled egg
   c. tortilla soup

3.53 State the temperature on the Fahrenheit thermometer and convert it to Celsius. (3.3)

3.54 State the temperature on the Celsius thermometer and convert to Fahrenheit. (3.3)

3.55 Compost can be made at home from grass clippings, some kitchen scraps, and dry leaves. As microbes break down organic matter, heat is generated and the compost can reach a temperature of 155 °F, which kills most pathogens. What is this temperature in degrees Celsius? In kelvins? (3.3)

3.56 In a chemical reaction, a mixture of salicylic acid and methanol was heated at 90 °C for two hours. After completion of the reaction, the mixture is cooled to 30 °C. What is this temperature in degrees Fahrenheit? In kelvins? (3.3)

3.57 Calculate the energy to heat two cubes (gold and aluminum) each with a volume of 20.0 cm³ from 25 °C to 35 °C. Refer to Tables 2.9 and 3.11. (3.6)

3.58 Calculate the energy needed, in kilojoules, to heat two cubes (iron and copper), each with a volume of 100.0 cm³ from 25 °C to 100 °C. Refer to Tables 2.9 and 3.7. Based on your calculations, explain the advantage of using copper in cookware. (3.5)

**Applications**

3.59 A 70.0-kg person had a quarter-pound cheeseburger, french fries, and a chocolate shake. (3.6)

<table>
<thead>
<tr>
<th>Item</th>
<th>Carbohydrate (g)</th>
<th>Fat (g)</th>
<th>Protein (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cheeseburger</td>
<td>46</td>
<td>40.47</td>
<td></td>
</tr>
<tr>
<td>French fries</td>
<td>47</td>
<td>16</td>
<td>4</td>
</tr>
<tr>
<td>Chocolate shake</td>
<td>76</td>
<td>10.10</td>
<td></td>
</tr>
</tbody>
</table>

a. Using Table 3.8, calculate the total kilocalories for each food type in this meal (round off the kilocalories to the tens place).
b. Determine the total kilocalories for the meal (round off to the tens place).
c. Using Table 3.11, determine the number of hours of sleep needed to burn off the kilocalories in this meal.
d. Using Table 3.11, determine the number of hours of running needed to burn off the kilocalories in this meal.

3.60 Your friend, who has a mass of 70.0 kg, has a slice of pizza, a cola soft drink, and ice cream. (3.6)

<table>
<thead>
<tr>
<th>Item</th>
<th>Carbohydrate (g)</th>
<th>Fat (g)</th>
<th>Protein (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pizza</td>
<td>29</td>
<td>10.13</td>
<td></td>
</tr>
<tr>
<td>Cola</td>
<td>51</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Ice cream</td>
<td>44</td>
<td>28.08</td>
<td>8</td>
</tr>
</tbody>
</table>

a. Using Table 3.8, calculate the total kilocalories for each food type in this meal (round off the kilocalories to the tens place).
b. Determine the total kilocalories for the meal (round off to the tens place).
c. Using Table 3.11, determine the number of hours of sitting needed to burn off the kilocalories in this meal.
d. Using Table 3.11, determine the number of hours of swimming needed to burn off the kilocalories in this meal.
3.61 Classify each of the following as an element, a compound, or a mixture: (3.1)
   a. iron filings in concrete
   b. sodium carbonate in soap
   c. air in the atmosphere

3.62 Classify each of the following as an element, a compound, or a mixture: (3.1)
   a. copper (Cu) in wiring
   b. chromium oxide in paint
   c. disinfectant in mouthwash

3.63 Classify each of the following mixtures as homogeneous or heterogeneous: (3.1)
   a. powdered iron and powdered sulfur
   b. vinegar
   c. toothpaste

3.64 Classify each of the following mixtures as homogeneous or heterogeneous: (3.1)
   a. dishwashing detergent
   b. ice in soda
   c. vinegar and oil mixture

3.65 Identify each of the following as solid, liquid, or gas: (3.2)
   a. calcium tablets in a bottle
   b. alcohol in water
   c. compressed air in a spray bottle
   d. dust in the air
   e. apple juice in a bottle

3.66 Identify each of the following as solid, liquid, or gas: (3.2)
   a. pencil in a box
   b. oil paint in a container
   c. ice cubes in a bowl
   d. salad dressing in a bowl
   e. air filled in a balloon

3.67 Identify each of the following as a physical or chemical property: (3.2)
   a. Mercury has a silvery shiny appearance.
   b. Mercury is a liquid in room temperature.
   c. The volume of mercury expands upon heating, therefore it is used as a thermometric liquid.
   d. Mercury reacts with yellow sulfur to form a red solid.

3.68 Identify each of the following as a physical or chemical property: (3.2)
   a. burning of gasoline to give CO₂ and water vapor
   b. souring of milk
   c. magnetizing a compass needle
   d. dicing potatoes into small pieces

3.69 Identify each of the following as a physical or chemical change: (3.2)
   a. Apples turn brown after cutting.
   b. A fertilized egg turns into an embryo.
   c. Damp clothes are dried under the sun.
   d. Egg white is set upon heating.

3.70 Identify each of the following as a physical or chemical change: (3.2)
   a. Aspirin tablets are broken in half.
   b. Carrots are grated for use in a salad.
   c. Malt undergoes fermentation to make beer.
   d. A copper pipe reacts with air and turns green.

3.71 Calculate each of the following temperatures in degrees Celsius and kelvins: (3.3)
   a. The highest recorded temperature in the continental United States was −69.7 °F in Rodgers Pass, Montana, on January 20, 1954.
   b. The lowest recorded temperature in the continental United States was −69.7 °F in Rodgers Pass, Montana, on January 20, 1954.

3.72 Calculate each of the following temperatures in kelvins and degrees Fahrenheit: (3.3)
   a. The highest recorded temperature in the world was 58.0 °C in El Azizia, Libya, on September 13, 1922.
   b. The lowest recorded temperature in the world was −89.2 °C in Vostok, Antarctica, on July 21, 1983.

3.73 What is −20 °F in degrees Celsius and in kelvins? (3.3)

3.74 On a hot sunny day, you get out of the swimming pool and sit in a metal chair, which is very hot. Would you predict that the specific heat of the metal is higher or lower than that of water? Explain. (3.5)

3.75 On a hot day, the beach sand gets hot but the water stays cool. Would you predict that the specific heat of sand is higher or lower than that of water? Explain. (3.5)

3.76 A large bottle containing 883 g of water at 4 °C is removed from the refrigerator. How many kilojoules are absorbed to warm the water to room temperature of 22 °C? (3.5)

3.77 A 0.50-g sample of vegetable oil is placed in a calorimeter. When the sample is burned, 18.9 kJ is given off. What is the energy value, in kcal/g, for the oil? (3.6)

3.78 A 1.3-g sample of rice is placed in a calorimeter. When the sample is burned, 22.00 kJ is given off. What is the energy value, in kcal/g, for the rice? (3.6)

3.79 When a substance with a mass of 45.6 g absorbs 575 J, its temperature increases from 21 °C to 35 °C. Calculate the specific heat, in J/g °C, of the substance. Identify the substance from Table 3.7. (3.5)

3.80 When a substance with a mass of 8.36 g absorbs 378 J, its temperature increases from 17.2 °C to 35.6 °C. Calculate the specific heat, in J/g °C, of the substance. Identify the substance from Table 3.7. (3.5)

3.81 Use the heat equation to calculate the energy, in joules, for each of the following (see Table 3.7): (3.5)
   a. lost when 15.0 g of ethanol, C₂H₅OH, cools from 60.5 °C to −42.0 °C
   b. lost when 125 g of iron (Fe) cools from 118 °C to 55 °C

3.82 Use the heat equation to calculate the energy, in joules, for each of the following (see Table 3.7): (3.5)
   a. lost when 75.0 g of water, H₂O, cools from 86.4 °C to 2.1 °C
   b. lost when 18.0 g of gold (Au) cools from 224 °C to 118 °C

3.83 Calculate the mass, in grams, for each of the following using Table 3.7: (3.5)
   a. a sample of aluminum (Al) that absorbs 8.80 kJ when heated from 12.5 °C to 26.8 °C
   b. a sample of titanium (Ti) that loses 14 200 J when it cools from 185 °C to 42 °C
3.84 Calculate the mass, in grams, for each of the following using Table 3.7: (3.5)
   a. a sample of silver (Ag) that absorbs 1650 J when its temperature increases from 65 °C to 187 °C
   b. an iron (Fe) bar that loses 2.52 kJ when its temperature decreases from 252 °C to 75 °C

3.85 Calculate the change in temperature (∆T) for each of the following using Table 3.7: (3.5)
   a. 85.0 g of gold (Au) that absorbs 7680 J
   b. 50.0 g of copper (Cu) that absorbs 6756 kJ

3.86 Calculate the change in temperature (∆T) for each of the following using Table 3.7: (3.5)
   a. 0.650 kg of water, H2O, that loses 5.48 kJ
   b. 35.0 g of silver (Ag) that loses 472 J

**CHALLENGE QUESTIONS**

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

3.91 When a 0.66-g sample of olive oil is burned in a calorimeter, the heat released increases the temperature of 370 g of water from 22.7 °C to 38.8 °C. What is the energy value, in kcal/g, for the olive oil? (3.5, 3.6)

3.92 When a 0.47-g sample of brown sugar is burned in a calorimeter, the heat released increases the temperature of 260 g of water from 21.6 °C to 28.9 °C. What is the energy value, in kcal/g, for the brown sugar? (3.5, 3.6)

3.93 In a large building, oil is used in a steam boiler heating system. The combustion of 1.0 lb of oil provides 2.4 × 10^7 J. How many kilograms of oil are needed to heat 150 kg of water from 22 °C to 100 °C? (3.4, 3.5)

3.94 When 1.0 g of gasoline burns, it releases 11 kcal of heat. The density of gasoline is 0.74 g/mL. (3.4, 3.5)
   a. How many megajoules are released when 1.0 gal of gasoline burns?
   b. If a television requires 150 kJ/h to run, how many hours can the television run on the energy provided by 1.0 gal of gasoline?

3.95 A 70.0-g piece of copper metal at 54.0 °C is placed in 50.0 g of water at 26.0 °C. If the final temperature of the water and metal is 29.2 °C, what is the specific heat (J/g °C) of copper? (3.5)

3.96 A 125-g piece of metal is heated to 288 °C and dropped into 85.0 g of water at 12.0 °C. The metal and water come to the same temperature of 24.0 °C. What is the specific heat, in J/g °C, of the metal? (3.5)

3.97 A metal is thought to be silver or iron. When 8.0 g of the metal absorbs 23.50 J, its temperature rises by 6.5 °C. (3.6)
   a. What is the specific heat (J/g °C) of the metal?
   b. Would you identify the metal as silver or iron (see Table 3.11)?

3.98 A metal is thought to be gold or titanium. When 20 g of the metal absorbs 30.8 cal, its temperature rises by 50 °C. (3.6)
   a. What is the specific heat, in cal/g °C, of the metal?
   b. Would you identify the metal as gold or titanium (see Table 3.11)?

**ANSWERS**

Answers to Selected Questions and Problems

3.1 a. pure substance b. mixture
c. pure substance d. pure substance
e. mixture

3.3 a. element b. compound
c. element d. compound
e. compound

3.5 a. heterogeneous b. homogeneous
c. homogeneous d. heterogeneous
e. heterogeneous

3.7 a. gas b. gas c. solid
d. chemical e. chemical

3.9 a. physical b. chemical c. physical
d. chemical e. chemical

3.11 a. physical b. chemical c. physical
d. physical e. physical

3.13 a. chemical b. physical c. physical
d. chemical e. physical

3.15 In the United States, we still use the Fahrenheit temperature scale. In °F, normal body temperature is 98.6. On the Celsius scale, her temperature would be 37.7 °C, a mild fever.

3.17 a. 98.6 °F b. 18.5 °C
c. 246 K d. 335 K e. 46 °C

3.19 a. 41 °C b. No. The temperature is equivalent to 39 °C.

3.21 When the roller-coaster car is at the top of the ramp, it has its maximum potential energy. As it descends, potential energy changes to kinetic energy. At the bottom, all the energy is kinetic.
### 3.23
- a. potential
- b. kinetic
- c. potential
- d. potential

### 3.25
- a. 3.5 kcal
- b. 99.2 cal
- c. 120 J
- d. 1100 cal

### 3.27
- a. 3.5 kcal
- b. 99.2 cal
- c. 120 J
- d. 1100 cal

### 3.29
Copper, which has the lowest specific heat, would reach the highest temperature.

### 3.31
- a. 0.389 J/g °C
- b. 0.313 J/g °C

### 3.33
- a. 1380 J
- b. 1810 J

### 3.35
- a. 54.5 g
- b. 216 g

### 3.37
- a. 175 °C
- b. 11.3 °C

### 3.39
- a. 29.9 kcal
- b. 208 kcal

### 3.41
- a. 470 kJ
- b. 18 g
- c. 130 kcal
- d. 37 g

### 3.43
210 kcal, 880 kJ

### 3.45
640 kcal

### 3.47
- a. Breakfast 270 kcal; lunch 420 kcal; dinner 440 kcal
- b. 1130 kcal total
- c. Yes. Charles should be losing weight.
- d. 26 days

### 3.49
- a. compound, the molecules have a definite 2:1 ratio of atoms
- b. mixture, has two different kinds of atoms and molecules
- c. element, has a single kind of atom

### 3.51
- a. heterogeneous
- b. heterogeneous
- c. homogeneous

### 3.53
- a. 61.4 °F, 16.3 °C
- b. 68.3 °C, 341 K

### 3.55
Gold, 497.94 J or 118.88 cal; aluminum, 484.38 J or 115.56 cal

### 3.57
- a. carbohydrate, 680 kcal; fat, 590 kcal; protein, 240 kcal
- b. 1510 kcal
- c. 25 h
- d. 2.0 h

### 3.61
- a. element
- b. compound
- c. mixture

### 3.63
- a. heterogeneous
- b. homogeneous

### 3.65
- a. solid
- b. liquid

### 3.67
- a. physical
- b. physical
- c. physical

### 3.69
- a. chemical
- b. chemical

### 3.71
- a. 56.7 °C, 330 K
- b. −56.5 °C, 217 K

### 3.73
- a. 210 kcal, 880 kJ

### 3.75
The same amount of heat causes a greater temperature change in the sand than in the water; thus the sand must have a lower specific heat than that of water.
CI.1 Gold, one of the most sought-after metals in the world, has a density of 19.3 g/cm³, a melting point of 1064 °C, and a specific heat of 0.129 J/g °C. A gold nugget found in Alaska in 1998 weighs 20.17 lb. (2.4, 2.6, 2.7, 3.3, 3.5)

a. How many significant figures are in the measurement of the weight of the nugget?
b. Which is the mass of the nugget in kilograms?
c. If the nugget were pure gold, what would its volume be in cubic centimeters?
d. What is the melting point of gold in degrees Fahrenheit and kelvins?
e. How many kilocalories are required to raise the temperature of the nugget from 500 °C to 1064 °C?
f. If the price of gold is $45.98 per gram, what is the nugget worth in dollars?

CI.2 The mileage for a motorcycle with a fuel-tank capacity of 22 L is 35 mi/gal. (2.5, 2.6, 2.7, 3.4)

a. How long a trip, in kilometers, can be made on one full tank of gasoline?
b. If the price of gasoline is $3.82 per gallon, what would be the cost of fuel for the trip?
c. If the average speed during the trip is 44 mi/h, how many hours will it take to reach the destination?
d. If the density of gasoline is 0.74 g/mL, what is the mass, in grams, of the fuel in the tank?
e. When 1.00 g of gasoline burns, 47 kJ of energy is released. How many kilojoules are produced when the fuel in one full tank is burned?

CI.3 Answer the following questions for the water samples A and B shown in the diagrams: (3.1, 3.2, 3.5)

a. In which sample (A or B) does the water have its own shape?
b. Which diagram (1 or 2 or 3) represents the arrangement of particles in water sample A?
c. Which diagram (1 or 2 or 3) represents the arrangement of particles in water sample B?

Answer the following for diagrams 1, 2, and 3: (3.2, 3.3)
d. The state of matter indicated in diagram 1 is a __________; in diagram 2, it is a __________; and in diagram 3, it is a __________.
e. The motion of the particles is slowest in diagram __________.
f. The arrangement of particles is farthest apart in diagram __________.
g. The particles fill the volume of the container in diagram __________.
h. If the water in diagram 2 has a mass of 19 g and a temperature of 45 °C, how much heat, in kilojoules, is removed to cool the liquid to 0 °C?

CI.4 The label of a black cherry almond energy bar with a mass of 68 g lists the nutrition facts as 39 g of carbohydrate, 5 g of fat, and 10 g of protein. (2.5, 2.6, 3.4, 3.6)

a. Using the energy values for carbohydrates, fats, and proteins (see Table 3.8), what are the total kilocalories listed for a black cherry almond bar? (Round off answers for each food type to the tens place.)
b. What are the kilojoules for the black cherry almond bar? (Round off answers for each food type to the tens place.)
c. If you obtain 160 kJ, how many grams of the black cherry almond bar did you eat?
d. If you are walking and using energy at a rate of 840 kJ/h, how many minutes will you need to walk to expend the energy from two black cherry almond bars?
CI.5 In one box of nails, there are 75 iron nails weighing 0.250 lb. The density of iron is 7.86 g/cm³. The specific heat of iron is 0.452 J/g °C. The melting point of iron is 1535 °C. (2.5, 2.6, 2.7, 3.4, 3.5)

a. What is the volume, in cubic centimeters, of the iron nails in the box?
b. If 30 nails are added to a graduated cylinder containing 17.6 mL of water, what is the new level of water, in milliliters, in the cylinder?
c. How much heat, in joules, must be added to the nails in the box to raise their temperature from 16 °C to 125 °C?
d. How much heat, in joules, is required to heat one nail from 25 °C to its melting point?

CI.6 A hot tub is filled with 450 gal of water. (2.5, 2.6, 2.7, 3.3, 3.4, 3.5)

a. What is the volume of water, in liters, in the tub?
b. What is the mass, in kilograms, of water in the tub?
c. How many kilocalories are needed to heat the water from 62 °F to 105 °F?
d. If the hot-tub heater provides 5900 kJ/min, how long, in minutes, will it take to heat the water in the hot tub from 62 °F to 105 °F?

ANSWERS

CI.1 a. 4 significant figures
b. 9.147 kg
c. 474 cm³
d. 1947 °F; 1337 K
e. 159 kcal
f. $421 000

cI.3 a. B
b. A is represented by diagram 2.
c. B is represented by diagram 1.
d. solid; liquid; gas
e. diagram 1
f. diagram 3
g. diagram 3
h. 3.6 kJ

CI.5 a. 14.4 cm³
b. 23.4 mL
c. 5590 J
d. 1030 J
Farming involves much more than growing crops and raising animals. Farmers must understand how to perform chemical tests and how to apply fertilizer to soil and pesticides or herbicides to crops. Pesticides are chemicals used to kill insects that could destroy the crop, whereas herbicides are chemicals used to kill weeds that would compete for the crop’s water and nutrient supply. This requires knowledge of how these chemicals work, their safety, effectiveness, and their storage. In using this information, farmers are able to grow crops that produce a higher yield, greater nutritional value, and better taste.

**CAREER**

**Farmer**

John is preparing for the next growing season as he decides how much of each crop should be planted and their location on his farm. Part of this decision is determined by the quality of the soil, including the pH, the amount of moisture, and the nutrient content in the soil. He begins by sampling the soil and performing a few chemical tests on the soil. John determines that several of his fields need additional fertilizer before the crops can be planted. John considers several different types of fertilizers as each supplies different nutrients to the soil to help increase crop production. Plants need three basic elements for growth. These elements are potassium, nitrogen, and phosphorus. Potassium (K on the periodic table) is a metal, whereas nitrogen (N) and phosphorus (P) are nonmetals. Fertilizers may also contain several other elements including calcium (Ca), magnesium (Mg), and sulfur (S). John applies a fertilizer containing a mixture of all of these elements to his soil and plans to re-check the soil nutrient content in a few days.
4.1 Elements and Symbols

**LEARNING GOAL** Given the name of an element, write its correct symbol; from the symbol, write the correct name.

All matter is composed of elements, of which there are 118 different kinds. Of these, 88 elements occur naturally and make up all the substances in our world. Many elements are already familiar to you. Perhaps you use aluminum in the form of foil or drink soft drinks from aluminum cans. You may have a ring or necklace made of gold, silver, or perhaps platinum. If you play tennis or golf, then you may have noticed that your racket or clubs may be made from the elements titanium or carbon. In our bodies, calcium and phosphorus form the structure of bones and teeth, iron and copper are needed in the formation of red blood cells, and iodine is required for the proper functioning of the thyroid.

Elements are pure substances from which all other things are built. Elements cannot be broken down into simpler substances. Over the centuries, elements have been named for planets, mythological figures, colors, minerals, geographic locations, and famous people. Some sources of names of elements are listed in **TABLE 4.1**. A complete list of all the elements and their symbols are found on the inside front cover of this text.

**TABLE 4.1** Some Elements, Symbols, and Source of Names

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Source of Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Uranium</td>
<td>U</td>
<td>The planet Uranus</td>
</tr>
<tr>
<td>Titanium</td>
<td>Ti</td>
<td>Titans (mythology)</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl</td>
<td>Chloros: “greenish yellow” (Greek)</td>
</tr>
<tr>
<td>Iodine</td>
<td>I</td>
<td>Ioeides: “violet” (Greek)</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg</td>
<td>Magnesia, a mineral</td>
</tr>
<tr>
<td>Californium</td>
<td>Cf</td>
<td>California</td>
</tr>
<tr>
<td>Curium</td>
<td>Cm</td>
<td>Marie and Pierre Curie</td>
</tr>
<tr>
<td>Copernicium</td>
<td>Cn</td>
<td>Nicolaus Copernicus</td>
</tr>
</tbody>
</table>

**Chemical Symbols**

Chemical symbols are one- or two-letter abbreviations for the names of the elements. Only the first letter of an element’s symbol is capitalized. If the symbol has a second letter, it is lowercase so that we know when a different element is indicated. If two letters are capitalized, they represent the symbols of two different elements. For example, the element cobalt has the symbol Co. However, the two capital letters CO specify two elements, carbon (C) and oxygen (O).

Although most of the symbols use letters from the current names, some are derived from their ancient names. For example, Na, the symbol for sodium, comes from the Latin
What are the names and symbols of some elements that you encounter every day? 

**TABLE 4.2** Names and Symbols of Some Common Elements

<table>
<thead>
<tr>
<th>Name*</th>
<th>Symbol</th>
<th>Name*</th>
<th>Symbol</th>
<th>Name*</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminum</td>
<td>Al</td>
<td>Gallium</td>
<td>Ga</td>
<td>Oxygen</td>
<td>O</td>
</tr>
<tr>
<td>Argon</td>
<td>Ar</td>
<td>Gold (aurum)</td>
<td>Au</td>
<td>Phosphorus</td>
<td>P</td>
</tr>
<tr>
<td>Arsenic</td>
<td>As</td>
<td>Helium</td>
<td>He</td>
<td>Platinum</td>
<td>Pt</td>
</tr>
<tr>
<td>Barium</td>
<td>Ba</td>
<td>Hydrogen</td>
<td>H</td>
<td>Potassium (kalium)</td>
<td>K</td>
</tr>
<tr>
<td>Boron</td>
<td>B</td>
<td>Iodine</td>
<td>I</td>
<td>Radium</td>
<td>Ra</td>
</tr>
<tr>
<td>Bromine</td>
<td>Br</td>
<td>Iron (ferrum)</td>
<td>Fe</td>
<td>Silicon</td>
<td>Si</td>
</tr>
<tr>
<td>Cadmium</td>
<td>Cd</td>
<td>Lead (plumbum)</td>
<td>Pb</td>
<td>Silver (argentum)</td>
<td>Ag</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca</td>
<td>Lithium</td>
<td>Li</td>
<td>Sodium (natrium)</td>
<td>Na</td>
</tr>
<tr>
<td>Carbon</td>
<td>C</td>
<td>Magnesium</td>
<td>Mg</td>
<td>Strontium</td>
<td>Sr</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl</td>
<td>Manganese</td>
<td>Mn</td>
<td>Sulfur</td>
<td>S</td>
</tr>
<tr>
<td>Chromium</td>
<td>Cr</td>
<td>Mercury (hydrargyrum)</td>
<td>Hg</td>
<td>Tin (stannum)</td>
<td>Sn</td>
</tr>
<tr>
<td>Cobalt</td>
<td>Co</td>
<td>Neon</td>
<td>Ne</td>
<td>Titanium</td>
<td>Ti</td>
</tr>
<tr>
<td>Copper (cuprum)</td>
<td>Cu</td>
<td>Nickel</td>
<td>Ni</td>
<td>Uranium</td>
<td>U</td>
</tr>
<tr>
<td>Fluorine</td>
<td>F</td>
<td>Nitrogen</td>
<td>N</td>
<td>Zinc</td>
<td>Zn</td>
</tr>
</tbody>
</table>

*Names given in parentheses are ancient Latin or Greek words from which the symbols are derived.

The symbol for iron, Fe, is derived from the Latin name ferrum. TABLE 4.2 lists the names and symbols of some common elements. Learning their names and symbols will greatly help your learning of chemistry.

**SAMPLE PROBLEM 4.1** Names and Symbols of Chemical Elements

Complete the following table with the correct name or symbol of each element:

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>nickel</td>
<td>______</td>
</tr>
<tr>
<td>nitrogen</td>
<td>______</td>
</tr>
<tr>
<td>______</td>
<td>Zn</td>
</tr>
<tr>
<td>______</td>
<td>K</td>
</tr>
<tr>
<td>iron</td>
<td>______</td>
</tr>
</tbody>
</table>

**SOLUTION**

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>nickel</td>
<td>Ni</td>
</tr>
<tr>
<td>nitrogen</td>
<td>N</td>
</tr>
<tr>
<td>zinc</td>
<td>Zn</td>
</tr>
<tr>
<td>potassium</td>
<td>K</td>
</tr>
<tr>
<td>iron</td>
<td>Fe</td>
</tr>
</tbody>
</table>

**STUDY CHECK 4.1**

Write the chemical symbols for the elements silicon, sulfur, and silver.

**ANSWER**

Si, S, and Ag
Carbon has the symbol C. However, its atoms can be arranged in different ways to give several different substances. Two forms of carbon—diamond and graphite—have been known since prehistoric times. A diamond is transparent and harder than any other substance, whereas graphite is black and soft. In diamond, carbon atoms are arranged in a rigid structure. In graphite, carbon atoms are arranged in sheets that slide over each other. Graphite is used as pencil lead, as lubricants, and as carbon fibers for the manufacture of lightweight golf clubs and tennis rackets.

Two other forms of carbon have been discovered more recently. In the form called Buckminsterfullerene or buckyball (named after R. Buckminster Fuller, who popularized the geodesic dome), 60 carbon atoms are arranged as rings of five and six atoms to give a spherical, cage-like structure. When a fullerene structure is stretched out, it produces a cylinder with a diameter of only a few nanometers called a nanotube. Practical uses for buckyballs and nanotubes are not yet developed, but they are expected to find use in lightweight structural materials, heat conductors, computer parts, and medicine. Recent research has shown that carbon nanotubes (CNT) can carry many drug molecules that can be released once the CNT enter the targeted cells.

Mercury (Hg) is a silvery, shiny element that is a liquid at room temperature. Mercury can enter the body through inhaled mercury vapor, contact with the skin, or ingestion of foods or water contaminated with mercury. In the body, mercury destroys proteins and disrupts cell function. Long-term exposure to mercury can damage the brain and kidneys, cause mental retardation, and decrease physical development. Blood, urine, and hair samples are used to test for mercury.

In both freshwater and seawater, bacteria convert mercury into toxic methylmercury, which attacks the central nervous system. Because fish absorb methylmercury, we are exposed to mercury by consuming mercury-contaminated fish. As levels of mercury ingested from fish became a concern, the Food and Drug Administration set a maximum level of one part mercury per million parts seafood (1 ppm), which is the same as 1 mg of mercury in every
kilogram of seafood. Fish higher in the food chain, such as swordfish and shark, can have such high levels of mercury that the U.S. Environmental Protection Agency (EPA) recommends they be consumed no more than once a week.

One of the worst incidents of mercury poisoning occurred in Minamata and Niigata, Japan, in 1965. At that time, the ocean was polluted with high levels of mercury from industrial wastes. Because fish were a major food in the diet, more than 2000 people were affected with mercury poisoning and died or developed neural damage. In the United States between 1988 and 1997, the use of mercury decreased by 75% when the use of mercury was banned in paint and pesticides, and regulated in batteries and other products. Certain batteries and compact fluorescent light bulbs (CFL) contain mercury, and instructions for their safe disposal should be followed.

### QUESTIONS AND PROBLEMS

#### 4.1 Elements and Symbols

**LEARNING GOAL** Given the name of an element, write its correct symbol; from the symbol, write the correct name.

4.1 Write the symbols for the following elements:
- a. copper
- b. platinum
- c. calcium
- d. manganese
- e. iron
- f. barium
- g. lead
- h. strontium

4.2 Write the symbols for the following elements:
- a. oxygen
- b. lithium
- c. uranium
- d. titanium
- e. hydrogen
- f. chromium
- g. tin
- h. gold

#### Applications

4.3 Write the name for the symbol of each of the following elements essential in the body:
- a. C
- b. Cl
- c. I
- d. Se
- e. N
- f. S
- g. Zn
- h. Co

4.4 Write the name for the symbol of each of the following elements essential in the body:
- a. V
- b. P
- c. Na
- d. As
- e. Ca
- f. Mo
- g. Mg
- h. Si

4.5 Write the names for the elements in each of the following formulas of compounds used in medicine:
- a. table salt, NaCl
- b. plaster casts, CaSO₄
- c. Demerol, C₁₅H₂₂ClNO₂
- d. treatment of bipolar disorder, Li₂CO₃

4.6 Write the names for the elements in each of the following formulas of compounds used in medicine:
- a. salt substitute, KCl
- b. dental cement, Zn₃(PO₄)₂
- c. antacid, Mg(OH)₂
- d. contrast agent for X-ray, BaSO₄

#### 4.2 The Periodic Table

**LEARNING GOAL** Use the periodic table to identify the group and the period of an element; identify the element as a metal, a nonmetal, or a metalloid.

As more elements were discovered, it became necessary to organize them into some type of classification system. By the late 1800s, scientists recognized that certain elements looked alike and behaved much the same way. In 1869, a Russian chemist, Dmitri Mendeleev, arranged the 60 elements known at that time into groups with similar properties and placed them in order of increasing atomic masses. Today, this arrangement of 118 elements is known as the periodic table (see Figure 4.1).

#### Periods and Groups

Each horizontal row in the periodic table is a period (see Figure 4.2). The periods are counted from the top of the table as Periods 1 to 7. The first period contains two elements: hydrogen (H) and helium (He). The second period contains eight elements: lithium (Li), beryllium (Be), boron (B), carbon (C), nitrogen (N), oxygen (O), fluorine (F), and neon (Ne). The third period also contains eight elements beginning with sodium (Na) and ending with argon (Ar). The fourth period, which begins with potassium (K), and the fifth period, which begins with rubidium (Rb), have 18 elements each. The sixth period, which begins with cesium (Cs), has 32 elements. The seventh period, as of today, contains 32 elements, for a total of 118 elements.

Each vertical column on the periodic table contains a group (or family) of elements that have similar properties. A group number is written at the top of each vertical column (group) in the periodic table. For many years, the representative elements have had group...
Periodic Table of Elements

Representative elements

<table>
<thead>
<tr>
<th>Period</th>
<th>Group 1A</th>
<th>Group 2A</th>
<th>Group 3A</th>
<th>Group 4A</th>
<th>Group 5A</th>
<th>Group 6A</th>
<th>Group 7A</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H</td>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
</tr>
<tr>
<td>2</td>
<td>Na</td>
<td>Mg</td>
<td>Al</td>
<td>Si</td>
<td>P</td>
<td>S</td>
<td>Cl</td>
</tr>
<tr>
<td>3</td>
<td>K</td>
<td>Ca</td>
<td>Sc</td>
<td>Ti</td>
<td>V</td>
<td>Cr</td>
<td>Mn</td>
</tr>
<tr>
<td>4</td>
<td>Rb</td>
<td>Sr</td>
<td>Y</td>
<td>Zr</td>
<td>Nb</td>
<td>Mo</td>
<td>Tc</td>
</tr>
<tr>
<td>5</td>
<td>Cs</td>
<td>Ba</td>
<td>La</td>
<td>Hf</td>
<td>Ta</td>
<td>W</td>
<td>Re</td>
</tr>
<tr>
<td>6</td>
<td>Fr</td>
<td>Ra</td>
<td>Ac</td>
<td>Th</td>
<td>Pa</td>
<td>U</td>
<td>Np</td>
</tr>
<tr>
<td>7</td>
<td>Rf</td>
<td>Db</td>
<td>Sg</td>
<td>Bh</td>
<td>Hs</td>
<td>Mt</td>
<td>Ds</td>
</tr>
</tbody>
</table>

Periods are the elements in each horizontal row.

- **Alkali metals**
- **Alkaline earth metals**
- **Transition elements**
- **Lanthanides**
- **Actinides**
- **Metals**
- **Metalloids**
- **Nonmetals**

**FIGURE 4.1** On the periodic table, groups are the elements arranged as vertical columns, and periods are the elements in each horizontal row.

What is the symbol of the alkali metal in Period 3?

**FIGURE 4.2** On the periodic table, each vertical column represents a group of elements and each horizontal row of elements represents a period.

Are the elements Si, P, and S part of a group or a period?

---

numbers 1A to 8A. In the center of the periodic table is a block of elements known as the **transition elements**, which have had group numbers followed by the letter “B.” A newer system assigns numbers of 1 to 18 to all of the groups going left to right across the periodic table. Because both systems are currently in use, they are both shown on the periodic table in this text and are included in our discussions of elements and group numbers. The two rows of 14 elements called the **lanthanides** and **actinides** (or the inner transition elements), which are part of Periods 6 and 7, are placed at the bottom of the periodic table to allow it to fit on a page.
Names of Groups

Several groups in the periodic table have special names (see FIGURE 4.3). Group 1A (1) elements—lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr)—are a family of elements known as the alkali metals (see FIGURE 4.4). The elements within this group are soft, shiny metals that are good conductors of heat and electricity and have relatively low melting points. Alkali metals react vigorously with water and form white products when they combine with oxygen.

Although hydrogen (H) is at the top of Group 1A (1), it is not an alkali metal and has very different properties than the rest of the elements in this group. Thus, hydrogen is not included in the alkali metals.

The alkaline earth metals are found in Group 2A (2). They include the elements beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). The alkaline earth metals are shiny metals like those in Group 1A (1), but they are not as reactive.

The halogens are found on the right side of the periodic table in Group 7A (17). They include the elements fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At) (see FIGURE 4.5). The halogens, especially fluorine and chlorine, are highly reactive and form compounds with most of the elements.

The noble gases are found in Group 8A (18). They include helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). They are quite unreactive and are seldom found in combination with other elements.

Metals, Nonmetals, and Metalloids

Another feature of the periodic table is the heavy zigzag line that separates the elements into the metals and the nonmetals. Except for hydrogen, the metals are to the left of the line with the nonmetals to the right (see FIGURE 4.6).

In general, most metals are shiny solids, such as copper (Cu), gold (Au), and silver (Ag). Metals can be shaped into wires (ductile) or hammered into a flat sheet (malleable). Metals are good conductors of heat and electricity. They usually melt at higher temperatures than nonmetals. All the metals are solids at room temperature, except for mercury (Hg), which is a liquid.

Nonmetals are not especially shiny, ductile, or malleable, and they are often poor conductors of heat and electricity. They typically have low melting points and low densities. Some examples of nonmetals are hydrogen (H), carbon (C), nitrogen (N), oxygen (O), chlorine (Cl), and sulfur (S).
Except for aluminum, the elements located along the heavy line are metalloids: B, Si, Ge, As, Sb, Te, Po, and At. Metalloids are elements that exhibit some properties that are typical of the metals and other properties that are characteristic of the nonmetals. For example, they are better conductors of heat and electricity than the nonmetals, but not as good as the metals. The metalloids are semiconductors because they can be modified to function as conductors or insulators. Table 4.3 compares some characteristics of silver, a metal, with those of antimony, a metalloid, and sulfur, a nonmetal.

**TABLE 4.3 Some Characteristics of a Metal, a Metalloid, and a Nonmetal**

<table>
<thead>
<tr>
<th>Silver (Ag)</th>
<th>Antimony (Sb)</th>
<th>Sulfur (S)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metal</td>
<td>Metalloid</td>
<td>Nonmetal</td>
</tr>
<tr>
<td>Shiny</td>
<td>Blue-gray, shiny</td>
<td>Dull, yellow</td>
</tr>
<tr>
<td>Extremely ductile</td>
<td>Brittle</td>
<td>Brittle</td>
</tr>
<tr>
<td>Can be hammered into sheets (malleable)</td>
<td>Shatters when hammered</td>
<td>Shatters when hammered</td>
</tr>
<tr>
<td>Good conductor of heat and electricity</td>
<td>Poor conductor of heat and electricity</td>
<td>Poor conductor of heat and electricity, good insulator</td>
</tr>
<tr>
<td>Used in coins, jewelry, tableware</td>
<td>Used to harden lead, color glass and plastics</td>
<td>Used in gunpowder, rubber, fungicides</td>
</tr>
<tr>
<td>Density 10.5 g/mL</td>
<td>Density 6.7 g/mL</td>
<td>Density 2.1 g/mL</td>
</tr>
<tr>
<td>Melting point 962 °C</td>
<td>Melting point 630 °C</td>
<td>Melting point 113 °C</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 4.2 Metals, Nonmetals, and Metalloids**

Use the periodic table to classify each of the following elements by its group and period, group name (if any), and as a metal, a nonmetal, or a metalloid:

a. Na, important in nerve impulses, regulates blood pressure
b. I, needed to produce thyroid hormones
c. Si, needed for tendons and ligaments

**TRY IT FIRST**

**SOLUTION**

a. Na (sodium), Group 1A (1), Period 3, is an alkali metal.
b. I (iodine), Group 7A (17), Period 5, halogen, is a nonmetal.
c. Si (silicon), Group 4A (14), Period 3, is a metalloid.
CHEMISTRY LINK TO HEALTH

Elements Essential to Health

Of all the elements, only about 20 are essential for the well-being and survival of the human body. Of those, four elements—oxygen, carbon, hydrogen, and nitrogen—which are representative elements in Period 1 and Period 2 on the periodic table, make up 96% of our body mass. Most of the food in our daily diet provides these elements to maintain a healthy body. These elements are found in carbohydrates, fats, and proteins. Most of the hydrogen and oxygen is found in water, which makes up 55% to 60% of our body mass.

The macrominerals—Ca, P, Cl, S, Na, and Mg—are representative elements located in Period 3 and Period 4 of the periodic table. They are involved in the formation of bones and teeth, maintenance of heart and blood vessels, muscle contraction, nerve impulses, acid–base balance of body fluids, and regulation of cellular metabolism. The macrominerals are present in lower amounts than the major elements, so that smaller amounts are required in our daily diets.

The other essential elements, called microminerals or trace elements, are mostly transition elements in Period 4 along with Si in Period 3 and Mo and I in Period 5. They are present in the human body in very small amounts, some less than 100 mg. In recent years, the detection of such small amounts has improved so that researchers can more easily identify the roles of trace elements. Some trace elements such as arsenic, chromium, and selenium are toxic at high levels in the body but are still required by the body. Other elements, such as tin and nickel, are thought to be essential, but their metabolic role has not yet been determined. Some examples and the amounts present in a 60-kg person are listed in Table 4.4.

![Periodic Table](image-url)
### TABLE 4.4 Typical Amounts of Essential Elements in a 60-kg Adult

<table>
<thead>
<tr>
<th>Element</th>
<th>Quantity</th>
<th>Function</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Major Elements</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Oxygen (O)</td>
<td>39 kg</td>
<td>Building block of biomolecules and water (H₂O)</td>
</tr>
<tr>
<td>Carbon (C)</td>
<td>11 kg</td>
<td>Building block of organic and biomolecules</td>
</tr>
<tr>
<td>Hydrogen (H)</td>
<td>6 kg</td>
<td>Component of biomolecules, water (H₂O), regulates pH of body fluids, stomach acid (HCl)</td>
</tr>
<tr>
<td>Nitrogen (N)</td>
<td>2 kg</td>
<td>Component of proteins and nucleic acids</td>
</tr>
<tr>
<td><strong>Macrominerals</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Calcium (Ca)</td>
<td>1000 g</td>
<td>Needed for bones and teeth, muscle contraction, nerve impulses</td>
</tr>
<tr>
<td>Phosphorus (P)</td>
<td>600 g</td>
<td>Needed for bones and teeth, nucleic acids</td>
</tr>
<tr>
<td>Potassium (K)</td>
<td>120 g</td>
<td>Most prevalent positive ion (K⁺) in cells, muscle contraction, nerve impulses</td>
</tr>
<tr>
<td>Chlorine (Cl)</td>
<td>100 g</td>
<td>Most prevalent negative ion (Cl⁻) in fluids outside cells, stomach acid (HCl)</td>
</tr>
<tr>
<td>Sulfur (S)</td>
<td>86 g</td>
<td>Component of proteins, liver, vitamin B₁, insulin</td>
</tr>
<tr>
<td>Sodium (Na)</td>
<td>60 g</td>
<td>Most prevalent positive ion (Na⁺) in fluids outside cells, water balance, muscle contraction, nerve impulses</td>
</tr>
<tr>
<td>Magnesium (Mg)</td>
<td>36 g</td>
<td>Component of bones, required for metabolic reactions</td>
</tr>
<tr>
<td><strong>Microminerals (Trace Elements)</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Iron (Fe)</td>
<td>3600 mg</td>
<td>Component of oxygen carrier hemoglobin</td>
</tr>
<tr>
<td>Silicon (Si)</td>
<td>3000 mg</td>
<td>Needed for growth and maintenance of bones and teeth, tendons and ligaments, hair and skin</td>
</tr>
<tr>
<td>Zinc (Zn)</td>
<td>2000 mg</td>
<td>Needed for metabolic reactions in cells, DNA synthesis, growth of bones, teeth, connective tissue, immune system</td>
</tr>
<tr>
<td>Copper (Cu)</td>
<td>240 mg</td>
<td>Needed for blood vessels, blood pressure, immune system</td>
</tr>
<tr>
<td>Manganese (Mn)</td>
<td>60 mg</td>
<td>Needed for growth of bones, blood clotting, metabolic reactions</td>
</tr>
<tr>
<td>Iodine (I)</td>
<td>20 mg</td>
<td>Needed for proper thyroid function</td>
</tr>
<tr>
<td>Molybdenum (Mo)</td>
<td>12 mg</td>
<td>Needed to process Fe and N from food</td>
</tr>
<tr>
<td>Arsenic (As)</td>
<td>3 mg</td>
<td>Needed for growth and reproduction</td>
</tr>
<tr>
<td>Chromium (Cr)</td>
<td>3 mg</td>
<td>Needed for maintenance of blood sugar levels, synthesis of biomolecules</td>
</tr>
<tr>
<td>Cobalt (Co)</td>
<td>3 mg</td>
<td>Component of vitamin B₁₂, red blood cells</td>
</tr>
<tr>
<td>Selenium (Se)</td>
<td>2 mg</td>
<td>Used in the immune system, health of heart and pancreas</td>
</tr>
<tr>
<td>Vanadium (V)</td>
<td>2 mg</td>
<td>Needed in the formation of bones and teeth, energy from food</td>
</tr>
</tbody>
</table>

### QUESTIONS AND PROBLEMS

#### 4.2 The Periodic Table

**LEARNING GOAL** Use the periodic table to identify the group and the period of an element; identify the element as a metal, a nonmetal, or a metalloid.

4.7 Identify the group or period number described by each of the following:
   a. contains C, N, and O
   b. begins with helium
   c. contains the alkali metals
   d. ends with neon

4.8 Identify the group or period number described by each of the following:
   a. contains Na, K, and Rb
   b. begins with Be
   c. contains the noble gases
   d. contains B, N, and F

4.9 Give the symbol of the element described by each of the following:
   a. Group 4A (14), Period 2
   b. the noble gas in Period 1
   c. the alkali metal in Period 3
   d. Group 2A (2), Period 4
   e. Group 3A (13), Period 3

4.10 Give the symbol of the element described by each of the following:
   a. the alkaline earth metal in Period 2
   b. Group 5A (15), Period 3
   c. the noble gas in Period 4
   d. the halogen in Period 5
   e. Group 4A (14), Period 4
4.11 Identify each of the following elements as a metal, a nonmetal, or a metalloid:
   a. calcium
   b. sulfur
   c. a shiny element
   d. an element that is a gas at room temperature
   e. located in Group 8A (18)
   f. bromine
   g. boron
   h. silver

4.12 Identify each of the following elements as a metal, a nonmetal, or a metalloid:
   a. located in Group 2A (2)
   b. a good conductor of electricity
   c. chlorine
   d. arsenic
   e. an element that is not shiny
   f. oxygen
   g. nitrogen
   h. tin

4.13 Using Table 4.4, identify the function of each of the following in the body and classify each as an alkali metal, an alkaline earth metal, a transition element, or a halogen:
   a. Ca
   b. Fe
   c. K
   d. Cl

4.14 Using Table 4.4, identify the function of each of the following in the body and classify each as an alkali metal, an alkaline earth metal, a transition element, or a halogen:
   a. Mg
   b. Cu
   c. I
   d. Na

4.15 Using the Chemistry Link to Health: Elements Essential to Health, answer each of the following:
   a. What is a macromineral?
   b. What is the role of sulfur in the human body?
   c. How many grams of sulfur would be a typical amount in a 60-kg adult?

4.16 Using the Chemistry Link to Health: Elements Essential to Health, answer each of the following:
   a. What is a micromineral?
   b. What is the role of iodine in the human body?
   c. How many milligrams of iodine would be a typical amount in a 60-kg adult?

4.3 The Atom

LEARNING GOAL Describe the electrical charge and location in an atom for a proton, a neutron, and an electron.

All the elements listed on the periodic table are made up of atoms. An atom is the smallest particle of an element that retains the characteristics of that element. Imagine that you are tearing a piece of aluminum foil into smaller and smaller pieces. Now imagine that you have a microscopic piece so small that it cannot be divided any further. Then you would have a single atom of aluminum.

The concept of the atom is relatively recent. Although the Greek philosophers in 500 B.C.E. reasoned that everything must contain minute particles they called atomos, the idea of atoms did not become a scientific theory until 1808. Then, John Dalton (1766–1844) developed an atomic theory that proposed that atoms were responsible for the combinations of elements found in compounds.

Dalton’s Atomic Theory

1. All matter is made up of tiny particles called atoms.
2. All atoms of a given element are similar to one another and different from atoms of other elements.
3. Atoms of two or more different elements combine to form compounds. A particular compound is always made up of the same kinds of atoms and always has the same number of each kind of atom.
4. A chemical reaction involves the rearrangement, separation, or combination of atoms. Atoms are neither created nor destroyed during a chemical reaction.

Dalton’s atomic theory formed the basis of current atomic theory, although we have modified some of Dalton’s statements. We now know that atoms of the same element are not completely identical to each other and consist of even smaller particles. However, an atom is still the smallest particle that retains the properties of an element.

Although atoms are the building blocks of everything we see around us, we cannot see an atom or even a billion atoms with the naked eye. However, when billions and billions of atoms are packed together, the characteristics of each atom are added to those of...
the next until we can see the characteristics we associate with the element. For example, a small piece of the element gold consists of many, many gold atoms. A special kind of microscope called a *scanning tunneling microscope* (STM) produces images of individual atoms (see **Figure 4.7**).

**Electrical Charges in an Atom**

By the end of the 1800s, experiments with electricity showed that atoms were not solid spheres but were composed of even smaller bits of matter called *subatomic particles*, three of which are the *proton*, *neutron*, and *electron*. Two of these subatomic particles were discovered because they have electrical charges.

An electrical charge can be positive or negative. Experiments show that like charges repel or push away from each other. When you brush your hair on a dry day, electrical charges that are alike build up on the brush and in your hair. As a result, your hair flies away from the brush. However, opposite or unlike charges attract. The crackle of clothes taken from the clothes dryer indicates the presence of electrical charges. The clinginess of the clothing results from the attraction of opposite, unlike charges (see **Figure 4.8**).

**Structure of the Atom**

In 1897, J. J. Thomson, an English physicist, applied electricity to a glass tube, which produced streams of small particles called *cathode rays*. Because these rays were attracted to a positively charged electrode, Thomson realized that the particles in the rays must be negatively charged. In further experiments, these particles called *electrons* were found to be much smaller than the atom and to have extremely small masses. Because atoms are neutral, scientists soon discovered that atoms contained positively charged particles called *protons* that were much heavier than the electrons.

Thomson proposed a “plum-pudding” model for the atom in which the electrons and protons were randomly distributed in a positively charged cloud like “plums in a pudding” or “chocolate chips in a cookie.” In 1911, Ernest Rutherford worked with Thomson to test this model. In Rutherford’s experiment, positively charged particles were aimed at a thin sheet of gold foil (see **Figure 4.9**). If the Thomson model were correct, the particles would travel in straight paths through the gold foil. Rutherford was greatly surprised to find that some of the particles were deflected as they passed through the gold foil, and a few particles were deflected so much that they went back in the opposite direction. According to Rutherford, it was as though he had shot a cannonball at a piece of tissue paper, and it bounced back at him.

From his gold-foil experiments, Rutherford realized that the protons must be contained in a small, positively charged region at the center of the atom, which he called the *nucleus*. He proposed that the electrons in the atom occupy the space surrounding the nucleus through which most of the particles traveled undisturbed. Only the particles that

**FIGURE 4.7** Images of gold atoms are produced when magnified 16 million times by a scanning microscope. Why is a microscope with extremely high magnification needed to see these atoms?

**FIGURE 4.8** Like charges repel and unlike charges attract. Why are the electrons attracted to the protons in the nucleus of an atom?

**FIGURE 4.9** Thomson’s “plum-pudding” model had protons and electrons scattered throughout the atom.
Chapter 4  Atoms and Elements

Atoms and Elements

came near this dense, positive center were deflected. If an atom were the size of a football stadium, the nucleus would be about the size of a golf ball placed in the center of the field.

Scientists knew that the nucleus was heavier than the mass of the protons, so they looked for another subatomic particle. Eventually, they discovered that the nucleus also contained a particle that is neutral, which they called a neutron. Thus, the masses of the protons and neutrons in the nucleus determine the mass of an atom (see Figure 4.10).

![Interactive Video](Rutherford's Gold-Foil Experiment)

**FIGURE 4.9**  (a) Positive particles are aimed at a piece of gold foil. (b) Particles that come close to the atomic nuclei are deflected from their straight path.

- Why are some particles deflected whereas most pass through the gold foil undeflected?

**FIGURE 4.10**  In an atom, the protons and neutrons that make up almost all the mass are packed into the tiny volume of the nucleus. The rapidly moving electrons (negative charge) surround the nucleus and account for the large volume of the atom.

- Why can we say that the atom is mostly empty space?

**Mass of the Atom**

All the subatomic particles are extremely small compared with the things you see around you. One proton has a mass of $1.67 \times 10^{-24}$ g, and the neutron is about the same. However, the electron has a mass $9.11 \times 10^{-28}$ g, which is much less than the mass of either a proton or neutron. Because the masses of subatomic particles are so small, chemists use a very small unit of mass called an atomic mass unit (amu). An amu is defined as one-twelfth of the mass of a carbon atom, which has a nucleus containing six protons and six
neutrons. In biology, the atomic mass unit is called a *Dalton* (Da) in honor of John Dalton. On the amu scale, the proton and neutron each have a mass of about 1 amu. Because the electron mass is so small, it is usually ignored in atomic mass calculations. **TABLE 4.5** summarizes some information about the subatomic particles in an atom.

TABLE 4.5 Subatomic Particles in the Atom

<table>
<thead>
<tr>
<th>Particle</th>
<th>Symbol</th>
<th>Charge</th>
<th>Mass (amu)</th>
<th>Location in Atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>$p$ or $p^+$</td>
<td>1+</td>
<td>1.007</td>
<td>Nucleus</td>
</tr>
<tr>
<td>Neutron</td>
<td>$n$ or $n^0$</td>
<td>0</td>
<td>1.008</td>
<td>Nucleus</td>
</tr>
<tr>
<td>Electron</td>
<td>$e^-$</td>
<td>1−</td>
<td>0.00055</td>
<td>Outside nucleus</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 4.3 Subatomic Particles**

Indicate whether each of the following is true or false:

a. A proton is heavier than an electron.

b. An electron is attracted to a neutron.

c. The nucleus contains all the protons and neutrons of an atom.

**TRY IT FIRST**

**SOLUTION**

a. True

b. False; an electron is attracted to a proton.

c. True

**STUDY CHECK 4.3**

Is the following statement true or false?
The nucleus occupies a large volume in an atom.

**ANSWER**

False, most of the volume of the atom is outside the nucleus.

**QUESTIONS AND PROBLEMS**

4.3 The Atom

**LEARNING GOAL** Describe the electrical charge and location in an atom for a proton, a neutron, and an electron.

4.17 Identify each of the following as describing either a proton, a neutron, or an electron:

a. has the smallest mass

b. has a 1+ charge

c. is found outside the nucleus

d. is electrically neutral

4.18 Identify each of the following as describing either a proton, a neutron, or an electron:

a. has a mass about the same as a proton

b. is found in the nucleus

c. is attracted to the protons

d. has a 1− charge

4.19 What did Rutherford determine about the structure of the atom from his gold-foil experiment?

4.20 How did Thomson determine that the electrons have a negative charge?

4.21 Is each of the following statements true or false?

a. A proton and an electron have opposite charges.

b. The nucleus contains most of the mass of an atom.

c. Electrons repel each other.

d. A proton is attracted to a neutron.

4.22 Is each of the following statements true or false?

a. A proton is attracted to an electron.

b. A neutron has twice the mass of a proton.

c. Neutrons repel each other.

d. Electrons and neutrons have opposite charges.

4.23 On a dry day, your hair flies apart when you brush it. How would you explain this?

4.24 Sometimes clothes cling together when removed from a dryer. What kinds of charges are on the clothes?
4.4 Atomic Number and Mass Number

**LEARNING GOAL.** Given the atomic number and the mass number of an atom, state the number of protons, neutrons, and electrons.

All the atoms of the same element always have the same number of protons. This feature distinguishes atoms of one element from atoms of all the other elements.

### Atomic Number

The **atomic number** of an element is equal to the number of protons in every atom of that element. The atomic number is the whole number that appears above the symbol of each element on the periodic table.

\[
\text{Atomic number} = \text{number of protons in an atom}
\]

The periodic table on the inside front cover of this text shows the elements in order of atomic number from 1 to 118. We can use an atomic number to identify the number of protons in an atom of any element. For example, a lithium atom, with atomic number 3, has 3 protons. Every lithium atom has 3 and only 3 protons. Any atom with 3 protons is always a lithium atom. In the same way, we determine that a carbon atom, with atomic number 6, has 6 protons. Every carbon atom has 6 protons, and any atom with 6 protons is carbon; every copper atom, with atomic number 29, has 29 protons and any atom with 29 protons is copper.

An atom is electrically neutral. That means that the number of protons in an atom is equal to the number of electrons, which gives every atom an overall charge of zero. Thus, for every atom, the atomic number also gives the number of electrons.

### Mass Number

We now know that the protons and neutrons determine the mass of the nucleus. Thus, for a single atom, we assign a **mass number**, which is the total number of protons and neutrons in its nucleus. However, the mass number does not appear on the periodic table because it applies to single atoms only.

\[
\text{Mass number} = \text{number of protons} + \text{number of neutrons}
\]

For example, the nucleus of a single oxygen atom that contains 8 protons and 8 neutrons has a mass number of 16. An atom of iron that contains 26 protons and 32 neutrons has a mass number of 58.
If we are given the mass number of an atom and its atomic number, we can calculate the number of neutrons in its nucleus.

**Number of neutrons in a nucleus** = **mass number** − **number of protons**

For example, if we are given a mass number of 37 for an atom of chlorine (atomic number 17), we can calculate the number of neutrons in its nucleus.

**Number of neutrons** = 37 (mass number) − 17 (protons) = 20 neutrons

**TABLE 4.6** illustrates these relationships between atomic number, mass number, and the number of protons, neutrons, and electrons in examples of single atoms for different elements.

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>H</td>
<td>1</td>
<td>1</td>
<td>1</td>
<td>0</td>
<td>1</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N</td>
<td>7</td>
<td>14</td>
<td>7</td>
<td>7</td>
<td>7</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O</td>
<td>8</td>
<td>16</td>
<td>8</td>
<td>8</td>
<td>8</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl</td>
<td>17</td>
<td>37</td>
<td>17</td>
<td>20</td>
<td>17</td>
</tr>
<tr>
<td>Potassium</td>
<td>K</td>
<td>19</td>
<td>39</td>
<td>19</td>
<td>20</td>
<td>19</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe</td>
<td>26</td>
<td>58</td>
<td>26</td>
<td>32</td>
<td>26</td>
</tr>
<tr>
<td>Gold</td>
<td>Au</td>
<td>79</td>
<td>197</td>
<td>79</td>
<td>118</td>
<td>79</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 4.4  Calculating Numbers of Protons, Neutrons, and Electrons**

Zinc, a micromineral, is needed for metabolic reactions in cells, DNA synthesis, the growth of bones, teeth, and connective tissue, and the proper functioning of the immune system. For an atom of zinc that has a mass number of 68, determine the following:

a. the number of protons  
b. the number of neutrons  
c. the number of electrons

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zn (Zn), mass number 68</td>
<td>number of protons, number of neutrons, number of electrons</td>
<td>periodic table, atomic number</td>
</tr>
</tbody>
</table>

a. Zinc (Zn), with an atomic number of 30, has 30 protons.

b. The number of neutrons in this atom is found by subtracting number of protons (atomic number) from the mass number.

\[
\text{Mass number } - \text{ atomic number } = \text{ number of neutrons} \\
68 - 30 = 38
\]

c. Because the zinc atom is neutral, the number of electrons is equal to the number of protons. A zinc atom has 30 electrons.

**STUDY CHECK 4.4**

How many neutrons are in the nucleus of a bromine atom that has a mass number of 80?

**ANSWER**

45
QUESTIONS AND PROBLEMS

4.4 Atomic Number and Mass Number

LEARNING GOAL Given the atomic number and the mass number of an atom, state the number of protons, neutrons, and electrons.

4.25 Would you use the atomic number, mass number, or both to determine each of the following?
   a. number of protons in an atom
   b. number of neutrons in an atom
   c. number of particles in the nucleus
   d. number of electrons in a neutral atom

4.26 Identify the type of subatomic particles described by each of the following:
   a. atomic number
   b. mass number
   c. mass number − atomic number
   d. mass number + atomic number

4.27 Write the names and symbols for the elements with the following atomic numbers:
   a. 3
   b. 9
   c. 20
   d. 30
   e. 10
   f. 14
   g. 53
   h. 8

4.28 Write the names and symbols for the elements with the following atomic numbers:
   a. 1
   b. 11
   c. 19
   d. 82
   e. 35
   f. 47
   g. 15
   h. 2

4.29 How many protons and electrons are there in a neutral atom of each of the following elements?
   a. argon
   b. manganese
   c. iodine
   d. cadmium

4.30 How many protons and electrons are there in a neutral atom of each of the following elements?
   a. carbon
   b. fluorine
   c. tin
   d. nickel

Applications

4.31 Complete the following table for atoms of essential elements in the body:

<table>
<thead>
<tr>
<th>Name of the Element</th>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zn</td>
<td>66</td>
<td>66</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>12</td>
<td>12</td>
<td>12</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Potassium</td>
<td></td>
<td></td>
<td>16</td>
<td>15</td>
<td>26</td>
</tr>
<tr>
<td></td>
<td></td>
<td>56</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

4.32 Complete the following table for atoms of essential elements in the body:

<table>
<thead>
<tr>
<th>Name of the Element</th>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>N</td>
<td>15</td>
<td>15</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>53</td>
<td>72</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>14</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>29</td>
<td></td>
<td></td>
<td>65</td>
</tr>
</tbody>
</table>

4.5 Isotopes and Atomic Mass

LEARNING GOAL Determine the number of protons, neutrons, and electrons in one or more of the isotopes of an element; calculate the atomic mass of an element using the percent abundance and mass of its naturally occurring isotopes.

We have seen that all atoms of the same element have the same number of protons and electrons. However, the atoms of any one element are not entirely identical because the atoms of most elements have different numbers of neutrons. When a sample of an element consists of two or more atoms with differing numbers of neutrons, those atoms are called isotopes.

Atoms and Isotopes

Isotopes are atoms of the same element that have the same atomic number but different numbers of neutrons. For example, all atoms of the element magnesium (Mg) have an atomic number of 12. Thus, every magnesium atom always has 12 protons. However, some naturally occurring magnesium atoms have 12 neutrons, others have 13 neutrons, and still others have 14 neutrons. The different numbers of neutrons give the magnesium atoms different mass numbers but do not change their chemical behavior. The three isotopes of magnesium have the same atomic number but different mass numbers.

To distinguish between the different isotopes of an element, we write an atomic symbol for a particular isotope that indicates the mass number in the upper left corner and the atomic number in the lower left corner.

An isotope may be referred to by its name or symbol, followed by its mass number, such as magnesium-24 or Mg-24. Magnesium has three naturally occurring isotopes, as
4.5 Isotopes and Atomic Mass

In a large sample of naturally occurring magnesium atoms, each type of isotope can be present as a low percentage or a high percentage. For example, the Mg-24 isotope makes up almost 80% of the total sample, whereas Mg-25 and Mg-26 each make up only about 10% of the total number of magnesium atoms.

<table>
<thead>
<tr>
<th>Atomic Symbol</th>
<th>Mass Number</th>
<th>Atomic Number</th>
<th>Percent Abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg-24</td>
<td>24</td>
<td>12</td>
<td>78.70</td>
</tr>
<tr>
<td>Mg-25</td>
<td>25</td>
<td>12</td>
<td>10.13</td>
</tr>
<tr>
<td>Mg-26</td>
<td>26</td>
<td>12</td>
<td>11.17</td>
</tr>
</tbody>
</table>

The nuclei of three naturally occurring magnesium isotopes have the same number of protons but different numbers of neutrons.

### SAMPLE PROBLEM 4.5 Identifying Protons and Neutrons in Isotopes

Chromium, a micromineral needed for maintenance of blood sugar levels, has four naturally occurring isotopes. Calculate the number of protons and number of neutrons in each of the following isotopes:

a. \(^{50}\text{Cr}\)

b. \(^{52}\text{Cr}\)

c. \(^{53}\text{Cr}\)

d. \(^{54}\text{Cr}\)

#### TRY IT FIRST

#### SOLUTION

In the atomic symbol, the mass number is shown in the upper left corner of the symbol, and the atomic number is shown in the lower left corner of the symbol. Thus, each isotope of Cr, atomic number 24, has 24 protons. The number of neutrons is found by subtracting the number of protons (24) from the mass number of each isotope.

#### STUDY CHECK 4.5

Vanadium is a micromineral needed in the formation of bones and teeth. Write the atomic symbol for the single naturally occurring isotope of vanadium, which has 27 neutrons.

#### ANSWER

\(^{50}\text{V}\)
Atomic Mass

In laboratory work, a chemist generally uses samples with many atoms that contain all the different atoms or isotopes of an element. Because each isotope has a different mass, chemists have calculated an atomic mass for an “average atom,” which is a weighted average of the masses of all the naturally occurring isotopes of that element. On the periodic table, the atomic mass is the number including decimal places that is given below the symbol of each element. Most elements consist of two or more isotopes, which is one reason that the atomic masses on the periodic table are seldom whole numbers.

Weighted Average Analogy

To understand how the atomic mass as a weighted average for a group of isotopes is calculated, we will use an analogy of bowling balls with different weights. Suppose that a bowling alley has ordered five bowling balls that weigh 8 lb each and 20 bowling balls that weigh 14 lb each. In this group of bowling balls, there are more 14-lb balls than 8-lb balls. The abundance of the 14-lb balls is 80% (20/25), and the abundance of the 8-lb balls is 20% (5/25). Now we can calculate a weighted average for the “average” bowling ball using the weight and percent abundance of the two types of bowling balls.

<table>
<thead>
<tr>
<th>Item</th>
<th>Weight</th>
<th>Percent Abundance</th>
<th>Weight from Each Type</th>
</tr>
</thead>
<tbody>
<tr>
<td>14-lb Bowling balls</td>
<td>14 lb</td>
<td>× 80 / 100</td>
<td>= 11.2 lb</td>
</tr>
<tr>
<td>8-lb Bowling balls</td>
<td>8 lb</td>
<td>× 20 / 100</td>
<td>= 1.6 lb</td>
</tr>
</tbody>
</table>

Weighted average of a bowling ball = 12.8 lb

“Atomic mass” of a bowling ball = 12.8 lb

Calculating Atomic Mass

To calculate the atomic mass of an element, we need to know the percentage abundance and the mass of each isotope, both of which must be determined experimentally. For example, a large sample of naturally occurring chlorine atoms consists of 75.76% of $^{35}$Cl atoms and 24.24% of $^{37}$Cl atoms. The $^{35}$Cl isotope has a mass of 34.97 amu and the $^{37}$Cl isotope has a mass of 36.97 amu.

Atomic mass of Cl = mass of $^{35}$Cl × $^{35}$Cl% / 100% + mass of $^{37}$Cl × $^{37}$Cl% / 100%

amu from $^{35}$Cl amu from $^{37}$Cl

<table>
<thead>
<tr>
<th>Atomic Symbol</th>
<th>Mass (amu)</th>
<th>×</th>
<th>Abundance (%)</th>
<th>=</th>
<th>Contribution to Average Cl Atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{35}$Cl</td>
<td>34.97</td>
<td>×</td>
<td>75.76</td>
<td>=</td>
<td>26.49 amu</td>
</tr>
<tr>
<td>$^{37}$Cl</td>
<td>36.97</td>
<td>×</td>
<td>24.24</td>
<td>=</td>
<td>8.962 amu</td>
</tr>
</tbody>
</table>

Atomic mass of Cl = 35.45 amu (weighted average mass)

The atomic mass of 35.45 amu is the weighted average mass of a sample of Cl atoms, although no individual Cl atom actually has this mass. An atomic mass of 35.45, which is closer to the mass number of Cl-35, also indicates there is a higher percentage of $^{35}$Cl atoms in the chlorine sample. In fact, there are about three atoms of $^{35}$Cl for every one atom of $^{37}$Cl in a sample of chlorine atoms.
TABLE 4.8 lists the naturally occurring isotopes of some selected elements and their atomic masses along with their most prevalent isotopes.

**TABLE 4.8 The Atomic Mass of Some Elements**

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Symbols</th>
<th>Atomic Mass (Weighted Average)</th>
<th>Most Prevalent Isotope</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>$^6$Li, $^7$Li</td>
<td>6.941 amu</td>
<td>$^7$Li</td>
</tr>
<tr>
<td>Carbon</td>
<td>$^{12}$C, $^{13}$C, $^{14}$C</td>
<td>12.01 amu</td>
<td>$^{12}$C</td>
</tr>
<tr>
<td>Oxygen</td>
<td>$^{16}$O, $^{17}$O, $^{18}$O</td>
<td>16.00 amu</td>
<td>$^{16}$O</td>
</tr>
<tr>
<td>Fluorine</td>
<td>$^{19}$F</td>
<td>19.00 amu</td>
<td>$^{19}$F</td>
</tr>
<tr>
<td>Sulfur</td>
<td>$^{32}$S, $^{33}$S, $^{34}$S, $^{36}$S</td>
<td>32.07 amu</td>
<td>$^{32}$S</td>
</tr>
<tr>
<td>Potassium</td>
<td>$^{39}$K, $^{40}$K, $^{41}$K</td>
<td>39.10 amu</td>
<td>$^{39}$K</td>
</tr>
<tr>
<td>Copper</td>
<td>$^{63}$Cu, $^{65}$Cu</td>
<td>63.55 amu</td>
<td>$^{63}$Cu</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 4.6 Calculating Atomic Mass**

Magnesium is a macromineral, which is a component of bone and needed for metabolic reactions. Using Table 4.7, calculate the atomic mass for magnesium using the weighted average mass method.

**TRY IT FIRST**

**SOLUTION**

**STEP 1** Multiply the mass of each isotope by its percent abundance divided by 100.

<table>
<thead>
<tr>
<th>Atomic Symbol</th>
<th>Mass (amu)</th>
<th>Abundance (%)</th>
<th>Contribution to the Atomic Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{24}$Mg</td>
<td>23.99</td>
<td>$\frac{78.70}{100}$</td>
<td>= 18.88 amu</td>
</tr>
<tr>
<td>$^{25}$Mg</td>
<td>24.99</td>
<td>$\frac{10.13}{100}$</td>
<td>= 2.531 amu</td>
</tr>
<tr>
<td>$^{26}$Mg</td>
<td>25.98</td>
<td>$\frac{11.17}{100}$</td>
<td>= 2.902 amu</td>
</tr>
</tbody>
</table>

**STEP 2** Add the contribution of each isotope to obtain the atomic mass.

Atomic mass of Mg = 18.88 amu + 2.531 amu + 2.902 amu

= 24.31 amu (weighted average mass)

**STUDY CHECK 4.6**

There are two naturally occurring isotopes of boron. The isotope $^{10}$B has a mass of 10.01 amu with an abundance of 19.80%, and the isotope $^{11}$B has a mass of 11.01 amu with an abundance of 80.20%. Calculate the atomic mass for boron using the weighted average mass method.

**ANSWER**

10.81 amu
4.33 What are the number of protons, neutrons, and electrons in the following isotopes?
   a. $^{89}_{38}\text{Sr}$  
   b. $^{52}_{24}\text{Cr}$  
   c. $^{34}_{16}\text{S}$  
   d. $^{81}_{35}\text{Br}$

4.34 What are the number of protons, neutrons, and electrons in the following isotopes?
   a. $^1_1\text{H}$  
   b. $^{14}_{7}\text{N}$  
   c. $^{26}_{14}\text{Si}$  
   d. $^{70}_{30}\text{Zn}$

4.35 Write the atomic symbol for the isotope with each of the following characteristics:
   a. 15 protons and 16 neutrons  
   b. 35 protons and 45 neutrons  
   c. 50 electrons and 72 neutrons  
   d. a chlorine atom with 18 neutrons  
   e. a mercury atom with 122 neutrons

4.36 Write the atomic symbol for the isotope with each of the following characteristics:
   a. an oxygen atom with 10 neutrons  
   b. 4 protons and 5 neutrons  
   c. 25 electrons and 28 neutrons  
   d. a mass number of 24 and 13 neutrons  
   e. a nickel atom with 32 neutrons

4.37 Argon has three naturally occurring isotopes, with mass numbers 36, 38, and 40.
   a. Write the atomic symbol for each of these atoms.  
   b. How are these isotopes alike?  
   c. How are they different?  
   d. Why is the atomic mass of argon listed on the periodic table not a whole number?  
   e. Which isotope is the most prevalent in a sample of argon?

4.38 Strontium has four naturally occurring isotopes, with mass numbers 84, 86, 87, and 88.
   a. Write the atomic symbol for each of these atoms.  
   b. How are these isotopes alike?  
   c. How are they different?  
   d. Why is the atomic mass of strontium listed on the periodic table not a whole number?  
   e. Which isotope is the most prevalent in a sample of strontium?

4.39 What is the difference between the mass of an isotope and the atomic mass of an element?

4.40 What is the difference between the mass number and the atomic mass of an element?

4.41 Two isotopes of gallium are naturally occurring, with $^{69}_{31}\text{Ga}$ at 60.11% (68.93 amu) and $^{71}_{31}\text{Ga}$ at 39.89% (70.92 amu). Calculate the atomic mass for gallium using the weighted average mass method.

4.42 Two isotopes of rubidium occur naturally, with $^{85}_{37}\text{Rb}$ at 72.17% (84.91 amu) and $^{87}_{37}\text{Rb}$ at 27.83% (86.91 amu). Calculate the atomic mass for rubidium using the weighted average mass method.

4.43 Copper consists of two isotopes, $^{63}_{29}\text{Cu}$ and $^{65}_{29}\text{Cu}$. If the atomic mass for copper on the periodic table is 63.55, are there more atoms of $^{63}_{29}\text{Cu}$ or $^{65}_{29}\text{Cu}$ in a sample of copper?

4.44 A fluorine sample consists of only one type of atom, $^{19}_{9}\text{F}$, which has a mass of 19.00 amu. How would the mass of a $^{19}_{9}\text{F}$ atom compare to the atomic mass listed on the periodic table?

4.45 There are two naturally occurring isotopes of thallium: $^{203}_{81}\text{Tl}$ and $^{205}_{81}\text{Tl}$. Use the atomic mass of thallium listed on the periodic table to identify the more prevalent isotope.

4.46 Zinc consists of five naturally occurring isotopes: $^{64}_{30}\text{Zn}$, $^{66}_{30}\text{Zn}$, $^{67}_{30}\text{Zn}$, $^{68}_{30}\text{Zn}$, and $^{70}_{30}\text{Zn}$. None of these isotopes has the atomic mass of $65.41$ listed for zinc on the periodic table. Explain.

Follow Up

IMPROVING CROP PRODUCTION

In plants, potassium (K) is needed for metabolic processes including the regulation of plant growth. Potassium is required by plants for protein synthesis, photosynthesis, enzymes, and ionic balance. Potassium-deficient potato plants may show purple or brown spots and plant, root, and seed growth is reduced. John has noticed that the leaves of his recent crop of potatoes had brown spots, the potatoes were undersized, and the crop yield was low.

Tests on soil samples showed that the potassium levels were below 100 ppm, which indicated that supplemental potassium was needed. John applied a fertilizer containing potassium chloride (KCl). To apply the correct amount of potassium, John needed to apply 170 kg of fertilizer per hectare.

Applications

4.47 a. What is the group number and name of the group that contains potassium?  
   b. Is potassium a metal, a nonmetal, or a metalloid?  
   c. How many protons are in an atom of potassium?  
   d. What is the most prevalent isotope of potassium (see Table 4.8)?
e. Potassium has three naturally occurring isotopes. They are K-39 (93.26%, 38.964 amu), K-40 (0.0117%, 39.964 amu), and K-41 (6.73%, 40.962 amu). Calculate the atomic mass of potassium from the naturally occurring isotopes and their abundance using the weighted average mass method.

4.48 a. How many neutrons are in K-41?
b. Write the electron configuration for potassium.
c. Which is the larger atom, K or Cs?
d. Which is the smallest atom, K, As, or Br?
e. If John’s potato farm has an area of 34.5 hectares, how many pounds of fertilizer does John need to use?
4.1 Elements and Symbols

**LEARNING GOAL** Given the name of an element, write its correct symbol; from the symbol, write the correct name.

- Elements are the primary substances of matter.
- Chemical symbols are one- or two-letter abbreviations of the names of the elements.

4.2 The Periodic Table

**LEARNING GOAL** Use the periodic table to identify the group and the period of an element; identify the element as a metal, a nonmetal, or a metalloid.

- The periodic table is an arrangement of the elements by increasing atomic number.
- A horizontal row is called a period.
- A vertical column on the periodic table containing elements with similar properties is called a group.
- Elements in Group 1A (1) are called the alkali metals; Group 2A (2), the alkaline earth metals; Group 7A (17), the halogens; and Group 8A (18), the noble gases.
- On the periodic table, metals are located on the left of the heavy zigzag line, and nonmetals are to the right of the heavy zigzag line.
- Except for aluminum, elements located along the heavy zigzag line are called metalloids.

4.3 The Atom

**LEARNING GOAL** Describe the electrical charge and location in an atom for a proton, a neutron, and an electron.

- An atom is the smallest particle that retains the characteristics of an element.
- Atoms are composed of three types of subatomic particles.
- Protons have a positive charge (+), electrons carry a negative charge (−), and neutrons are electrically neutral.
- The protons and neutrons are found in the tiny, dense nucleus; electrons are located outside the nucleus.

4.4 Atomic Number and Mass Number

**LEARNING GOAL** Given the atomic number and the mass number of an atom, state the number of protons, neutrons, and electrons.

- The atomic number gives the number of protons in all the atoms of the same element.
- In a neutral atom, the number of protons and electrons is equal.
- The mass number is the total number of protons and neutrons in an atom.

4.5 Isotopes and Atomic Mass

**LEARNING GOAL** Determine the number of protons, neutrons, and electrons in one or more of the isotopes of an element; calculate the atomic mass of an element using the percent abundance and mass of its naturally occurring isotopes.

- Atoms that have the same number of protons but different numbers of neutrons are called isotopes.
- The atomic mass of an element is the weighted average mass of all the isotopes in a naturally occurring sample of that element.

**KEY TERMS**

- **alkali metal** An element in Group 1A (1), except hydrogen.
- **alkaline earth metal** An element in Group 2A (2).
- **atom** The smallest particle of an element.
- **atomic mass** The weighted average mass of all the naturally occurring isotopes of an element.
- **atomic mass unit (amu)** A small mass unit used to describe the mass of extremely small particles such as atoms and subatomic particles; 1 amu is equal to one-twelfth the mass of a ¹²C atom.
- **atomic number** A number that is equal to the number of protons in an atom.
- **atomic symbol** An abbreviation used to indicate the mass number and atomic number of an isotope.
- **chemical symbol** An abbreviation that represents the name of an element.
- **electron** A negatively charged subatomic particle having a minute mass that is usually ignored in mass calculations; its symbol is e⁻.
group A vertical column in the periodic table that contains elements having similar physical and chemical properties.
group number A number that appears at the top of each vertical column (group) in the periodic table.
halogen An element in Group 7A (17).
isotope An atom that differs only in mass number from another atom of the same element. Isotopes have the same atomic number (number of protons), but different numbers of neutrons.
mass number The total number of protons and neutrons in the nucleus of an atom.
metal An element that is shiny, malleable, ductile, and a good conductor of heat and electricity. The metals are located to the left of the heavy zigzag line on the periodic table.
metalloid Elements with properties of both metals and nonmetals located along the heavy zigzag line on the periodic table.
neutron A neutral subatomic particle having a mass of about 1 amu and found in the nucleus of an atom; its symbol is $n$ or $n^0$.
noble gas An element in Group 8A (18) of the periodic table.
nonmetal An element with little or no luster that is a poor conductor of heat and electricity. The nonmetals are located to the right of the heavy zigzag line on the periodic table.
nucleus The compact, extremely dense center of an atom, containing the protons and neutrons of the atom.
period A horizontal row of elements in the periodic table.
periodic table An arrangement of elements by increasing atomic number such that elements having similar chemical behavior are grouped in vertical columns.
proton A positively charged subatomic particle having a mass of about 1 amu and found in the nucleus of an atom; its symbol is $p$ or $p^+$.
representative element An element in the first two columns on the left of the periodic table and the last six columns on the right that has a group number of 1A through 8A or 1, 2, and 13 through 18.
subatomic particle A particle within an atom; protons, neutrons, and electrons are subatomic particles.
transition element An element in the center of the periodic table that is designated with the letter “B” or the group number of 3 through 12.

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

**Writing Atomic Symbols for Isotopes (4.5)**

- Isotopes are atoms of the same element that have the same atomic number but different numbers of neutrons.
- An atomic symbol is written for a particular isotope, with its mass number (protons and neutrons) shown in the upper left corner and its atomic number (protons) shown in the lower left corner.

**Example:** Calculate the number of protons and neutrons in the cadmium isotope $^{112}\text{Cd}$.

**Answer:**

<table>
<thead>
<tr>
<th>Atomic Symbol</th>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{112}\text{Cd}$</td>
<td>number in lower left corner 48</td>
<td>number in upper left corner 112</td>
<td>equal to atomic number 48</td>
<td>equal to mass number $112 - 48 = 64$</td>
<td>equal to number of protons $48 - 36 = 12$</td>
</tr>
</tbody>
</table>
**UNDERSTANDING THE CONCEPTS**

The chapter sections to review are shown in parentheses at the end of each question.

4.49 According to Dalton’s atomic theory, which of the following are true or false? If false, correct the statement to make it true. (4.3)
   a. Atoms of an element are identical to atoms of other elements.
   b. Every element is made of atoms.
   c. Atoms of different elements combine to form compounds.
   d. In a chemical reaction, some atoms disappear and new atoms appear.

4.50 Use Rutherford’s gold-foil experiment to answer each of the following: (4.3)
   a. What did Rutherford expect to happen when he aimed particles at the gold foil?
   b. How did the results differ from what he expected?
   c. How did he use the results to propose a model of the atom?

4.51 Match the subatomic particles (1 to 3) to each of the descriptions below: (4.4)
   1. protons  2. neutrons  3. electrons
   a. atomic mass
   b. atomic number
   c. positive charge
   d. negative charge
   e. mass number – atomic number

4.52 Match the subatomic particles (1 to 3) to each of the descriptions below: (4.4)
   1. protons  2. neutrons  3. electrons
   a. mass number
   b. surround the nucleus
   c. in the nucleus
   d. charge of 0
   e. equal to number of electrons

4.53 Consider the following atoms in which X represents the chemical symbol of the element: (4.4, 4.5)

   \[ ^{16}_{8}O, ^{18}_{8}O, ^{17}_{8}O, ^{17}_{8}O, ^{18}_{8}O \]
   a. Which atoms have the same number of protons?
   b. Which atoms are isotopes? Of what element?
   c. Which atoms have the same mass number?

4.54 Consider the following atoms in which X represents the chemical symbol of the element: (4.4, 4.5)

   \[ ^{12}_{6}C, ^{13}_{6}C, ^{14}_{6}C, ^{15}_{6}C, ^{16}_{6}C \]
   a. Which atoms have the same number of protons?
   b. Which atoms are isotopes? Of what element?
   c. Which atoms have the same mass number?

4.55 Complete the following table for two of the naturally occurring isotopes of calcium, which is an important mineral for bone and teeth growth, muscle contraction, and nerve impulses: (4.4, 4.5)

<table>
<thead>
<tr>
<th>Atomic Symbol</th>
<th>( ^{42}_{20}Ca )</th>
<th>( ^{43}_{20}Ca )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atomic Number</td>
<td>20</td>
<td>20</td>
</tr>
<tr>
<td>Mass Number</td>
<td>42</td>
<td>43</td>
</tr>
<tr>
<td>Number of Protons</td>
<td>20</td>
<td>20</td>
</tr>
<tr>
<td>Number of Neutrons</td>
<td>22</td>
<td>23</td>
</tr>
<tr>
<td>Number of Electrons</td>
<td>20</td>
<td>20</td>
</tr>
</tbody>
</table>

4.56 Complete the following table for the three naturally occurring isotopes of silicon, the major component in computer chips: (4.4, 4.5)

<table>
<thead>
<tr>
<th>Atomic Symbol</th>
<th>( ^{28}_{14}Si )</th>
<th>( ^{29}_{14}Si )</th>
<th>( ^{30}_{14}Si )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atomic Number</td>
<td>14</td>
<td>14</td>
<td>14</td>
</tr>
<tr>
<td>Mass Number</td>
<td>28</td>
<td>29</td>
<td>30</td>
</tr>
<tr>
<td>Number of Protons</td>
<td>14</td>
<td>14</td>
<td>14</td>
</tr>
<tr>
<td>Number of Neutrons</td>
<td>14</td>
<td>15</td>
<td>16</td>
</tr>
<tr>
<td>Number of Electrons</td>
<td>14</td>
<td>14</td>
<td>14</td>
</tr>
</tbody>
</table>

4.57 For each representation of a nucleus A through E, write the atomic symbol, and identify which are isotopes. (4.4, 4.5)

- Proton
- Neutron

4.58 Identify the element represented by each nucleus A through E in problem 4.57 as a metal, a nonmetal, or a metalloid. (4.2)
ADDITIONAL QUESTIONS AND PROBLEMS

4.59 Why is Co the symbol for cobalt, not CO? (4.1)
4.60 Which of the following is correct? Write the correct symbol if needed. (4.1)
   a. copper, Cp
   b. silicon, Si
   c. iron, Fe
   d. fluorine, Fl
   e. potassium, P
   f. sodium, Na
   g. gold, Au
4.61 Give the group and period number for each of the following elements: (4.2)
   a. krypton
   b. bismuth
   c. gallium
4.62 Give the group and period number for each of the following elements: (4.2)
   a. strontium
   b. germanium
   c. sulfur
   d. fluorine
4.63 The following trace elements have been found to be crucial to the functions of the body. Indicate each as a metal, a nonmetal, or a metalloid. (4.2)
   a. silicon
   b. molybdenum
   c. sodium
   d. phosphorus
4.64 The following trace elements have been found to be crucial to the functions of the body. Indicate each as a metal, a nonmetal, or a metalloid. (4.2)
   a. iron
   b. vanadium
   c. calcium
   d. arsenic
4.65 Indicate if each of the following statements is true or false: (4.3)
   a. The proton is a negatively charged particle.
   b. The neutron is 2000 times as heavy as a proton.
   c. The atomic mass unit is based on a carbon atom with 6 protons and 6 neutrons.
   d. The nucleus is the largest part of the atom.
   e. The electrons are located outside the nucleus.
4.66 Indicate if each of the following statements is true or false: (4.3)
   a. The neutron is electrically neutral.

CHALLENGE QUESTIONS

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

4.71 Complete the following statements: (4.2, 4.4)
   a. The atomic number gives the number of _________ in the nucleus.
   b. In an atom, the number of electrons is equal to the number of _________.
   c. Sodium and potassium are examples of elements called _________.
4.72 Complete the following statements: (4.2, 4.4)
   a. The number of protons and neutrons in an atom is the _________ number.
   b. The elements in Group 7A (17) are called the _________.
   c. Elements that are shiny and conduct heat are called _________.
4.73 Provide the following information: (4.2, 4.4)
   a. the atomic number and symbol of the lightest alkali metal
   b. the atomic number and symbol of the heaviest noble gas
   c. the atomic mass and symbol of the alkaline earth metal in Period 3
   d. the atomic mass and symbol of the halogen with the fewest electrons
4.74 Provide the following information: (4.2, 4.4)
   a. the atomic number and symbol of the heaviest metalloid in Group 4A (14)
   b. the atomic number and symbol of the element in Group 5A (15), Period 6
   c. the atomic mass and symbol of the alkali metal in Period 4
   d. the metalloid in Group 3A (13)
4.75 Write the names and symbols of the elements with the following numbers of protons: (4.4)
   a. 27
   b. 32
   c. 11
   d. 48
   e. 62
   f. 73
4.76 With the help of the periodic table, write the numbers of protons in and the group positions of the following elements: (4.4)
   a. Bromine   b. Rhenium   c. Ruthenium
d. Zinc   e. Gallium   f. Lead

4.77 Give the number of protons and electrons in neutral atoms of each of the following: (4.4)
   a. Mn   b. phosphorus   c. Sr
d. Co   e. uranium

4.78 Give the number of protons and electrons in neutral atoms of each of the following: (4.4)
   a. chromium   b. Cs   c. copper
d. chlorine   e. Cd

4.79 The most abundant isotope of nickel is $^{58}_{28}$Ni. (4.4)
   a. How many protons, neutrons, and electrons are in $^{58}_{28}$Ni?
   b. What is the atomic symbol of another isotope of nickel with 35 neutrons?
   c. What is the name and symbol of an atom with the same mass number as in part b and 34 neutrons?

4.80 The most abundant isotope of strontium is $^{88}_{38}$Sr. (4.4)
   a. How many protons, neutrons, and electrons are in $^{88}_{38}$Sr?
   b. What is the atomic symbol of another isotope of strontium with 51 neutrons?
   c. What is the name and symbol of an atom with the same mass number as in part b and 50 neutrons?

4.81 Write the atomic symbol for each of the following: (4.5)
   a. an atom with only 1 proton
   b. an atom with 80 protons and 120 neutrons
   c. an osmium atom with mass number of 189
d. an atom with 24 electrons and 28 neutrons

4.82 Write the atomic symbol for each of the following: (4.5)
   a. an atom with 47 protons and mass number of 107
   b. an atom with 74 protons and 110 neutrons
   c. an atom with atomic number 70 and 100 neutrons
d. an atom with 50 electrons and 70 neutrons

4.83 Silicon has three naturally occurring isotopes: Si-28 that has a percent abundance of 92.23% and a mass of 27.977 amu, Si-29 that has a 4.68% abundance and a mass of 28.976 amu, and Si-30 that has a 3.09% abundance and a mass of 29.974 amu. Calculate the atomic mass for silicon using the weighted average mass method. (4.5)

4.84 Antimony (Sb) has two naturally occurring isotopes: Sb-121 that has a percent abundance of 57.21% and a mass of 120.90 amu, and Sb-123 that has a 42.79% abundance and a mass of 122.90 amu. Calculate the atomic mass for antimony using the weighted average mass method. (4.5)

4.85 The most prevalent isotope of gold is Au-197. (4.5)
   a. How many protons, neutrons, and electrons are in this isotope?
   b. What is the atomic symbol of another isotope of gold with 116 neutrons?
   c. What is the atomic symbol of an atom with an atomic number of 78 and 116 neutrons?

4.86 Cadmium, atomic number 48, consists of eight naturally occurring isotopes. Do you expect any of the isotopes to have the atomic mass listed on the periodic table for cadmium? Explain. (4.5)

4.87 Lead consists of four naturally occurring isotopes. Calculate the atomic mass for lead using the weighted average mass method. (4.5)

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass (amu)</th>
<th>Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{204}_{82}$Pb</td>
<td>204.0</td>
<td>1.40</td>
</tr>
<tr>
<td>$^{206}_{82}$Pb</td>
<td>206.0</td>
<td>24.10</td>
</tr>
<tr>
<td>$^{207}_{82}$Pb</td>
<td>207.0</td>
<td>22.10</td>
</tr>
<tr>
<td>$^{208}_{82}$Pb</td>
<td>208.0</td>
<td>52.40</td>
</tr>
</tbody>
</table>

4.88 Indium (In) has two naturally occurring isotopes: In-113 and In-115. In-113 has a 4.30% abundance and a mass of 112.9 amu, and In-115 has a 95.70% abundance and a mass of 114.9 amu. Calculate the atomic mass for indium using the weighted average mass method. (4.5)

4.89 If the diameter of a sodium atom is $3.14 \times 10^{-8}$ cm, how many sodium atoms would fit along a line exactly 1 in. long? (4.3)

4.90 A lead atom has a mass of $3.4 \times 10^{-22}$ g. How many lead atoms are in a cube of lead that has a volume of 2.00 cm$^3$ if the density of lead is 11.3 g/cm$^3$? (4.3)

ANSWERS

Answers to Selected Questions and Problems

4.1  a. Cu   b. Pt   c. Ca   d. Mn
e. Fe   f. Ba   g. Pb   h. Sr
4.3  a. carbon   b. chlorine   c. iodine
d. selenium   e. nitrogen   f. sulfur
g. zinc   h. cobalt
4.5  a. sodium, chlorine
   b. calcium, sulfur, oxygen
c. carbon, hydrogen, chlorine, nitrogen, oxygen
d. lithium, carbon, oxygen
4.7  a. Period 2
   b. Group 8A (18)
c. Group 1A (1)
d. Period 2
4.9  a. C   b. He   c. Na
d. Ca   e. Al
4.11 a. metal   b. nonmetal   c. metal
d. nonmetal   e. nonmetal   f. nonmetal
g. metalloid   h. metal
4.13 a. needed for bones and teeth, muscle contraction, nerve impulses; alkaline earth metal
   b. component of hemoglobin; transition element
c. muscle contraction, nerve impulses; alkali metal
d. found in fluids outside cells; halogen
4.15 a. A macromineral is an element essential to health, which is present in the body in amounts from 5 to 1000 g.
b. Sulfur is a component of proteins, liver, vitamin B$_1$, and insulin.
c. 86 g
4.17 a. electron   b. proton
c. electron   d. neutron
4.19 Rutherford determined that an atom contains a small, compact nucleus that is positively charged.

4.21 a. True  
   b. True  
   c. True  
   d. False; a proton is attracted to an electron

4.23 In the process of brushing hair, strands of hair become charged with like charges that repel each other.

4.25 a. atomic number  
   b. both  
   c. mass number  
   d. atomic number

4.27 a. lithium, Li  
   b. fluorine, F  
   c. calcium, Ca  
   d. zinc, Zn  
   e. neon, Ne  
   f. silicon, Si  
   g. iodine, I

4.29 a. 18 protons and 18 electrons  
   b. 25 protons and 25 electrons  
   c. 53 protons and 53 electrons  
   d. 48 protons and 48 electrons

4.31 a. 38 protons, 51 neutrons, 38 electrons  
   b. 24 protons, 28 neutrons, 24 electrons  
   c. 16 protons, 18 neutrons, 16 electrons  
   d. 35 protons, 46 neutrons, 35 electrons

4.33 a. 31P  
   b. 80Br  
   c. 122Sn  
   d. 35Cl  
   e. 202Hg

4.35 a. 36Ar  
   b. 38Ar  
   c. 40Ar  
   d. They all have the same number of protons and electrons.  
   e. They have different numbers of neutrons, which gives them different mass numbers.  
   f. The atomic mass of Ar listed on the periodic table is the weighted average atomic mass of all the naturally occurring isotopes.  
   g. The isotope Ar-40 is the most prevalent because its mass is closest to the atomic mass of Ar on the periodic table.

4.37 The mass of an isotope is the mass of an individual atom. The atomic mass is the weighted average of all the naturally occurring isotopes of that element.

4.41 69.72 amu

4.42 Since the atomic mass of copper is closer to 63 amu, there are more atoms of 63Cu.

4.43 Since the atomic mass of thallium is 204.4 amu, the most prevalent isotope is 205Tl.

4.45 a. Group 1A (1), alkali metals  
   b. metal  
   c. 19 protons  
   d. K-39  
   e. 39.10 amu

4.47 a. False; All atoms of a given element are different from atoms of other elements.  
   b. True  
   c. True  
   d. False; In a chemical reaction, atoms are neither created nor destroyed.

4.51 a. 1 + 2  
   b. 1  
   c. 1  
   d. 3  
   e. 2

4.53 a. 43X and 44X both have 20 protons.  
   b. 43X and 44X both are isotopes of calcium.  
   c. 45X and 46X have mass numbers of 18, and 47X and 48X have mass numbers of 42.

4.55

<table>
<thead>
<tr>
<th>Name of the Element</th>
<th>Atomic Symbol</th>
<th>42Ca</th>
<th>43Ca</th>
<th>44Ca</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zinc</td>
<td>Zn 30</td>
<td>66</td>
<td>30</td>
<td>36</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg 12</td>
<td>24</td>
<td>12</td>
<td>12</td>
</tr>
<tr>
<td>Potassium</td>
<td>K 19</td>
<td>39</td>
<td>19</td>
<td>20</td>
</tr>
<tr>
<td>Sulfur</td>
<td>S 16</td>
<td>31</td>
<td>16</td>
<td>15</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe 26</td>
<td>56</td>
<td>26</td>
<td>30</td>
</tr>
</tbody>
</table>

4.57 a. 9Be  
   b. 14B  
   c. 13C  
   d. 5B  
   e. 12C

4.59 The first letter of a symbol is a capital, but a second letter is lowercase. The symbol Co is for cobalt, but the symbols in CO are for carbon and oxygen.

4.61 a. Group 18 (8A), Period 4  
   b. Group 5A (15), Period 6  
   c. Group 3A (13), Period 4  
   d. Group 1 (1A), Period 6

4.63 a. metalloid  
   b. metal  
   c. metal  
   d. non-metal

4.65 a. false  
   b. false  
   c. true  
   d. false  
   e. true

4.67 a. 48 protons, 66 neutrons, 48 electrons  
   b. 43 protons, 55 neutrons, 43 electrons  
   c. 79 protons, 79 electrons  
   d. 86 protons, 86 electrons  
   e. 54 protons, 54 electrons

4.69

<table>
<thead>
<tr>
<th>Name of the Element</th>
<th>Atomic Symbol</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Selenium</td>
<td>34</td>
<td>46</td>
<td>34</td>
<td></td>
</tr>
<tr>
<td>Nickel</td>
<td>28</td>
<td>34</td>
<td>28</td>
<td></td>
</tr>
<tr>
<td>Magnesium</td>
<td>12</td>
<td>14</td>
<td>12</td>
<td></td>
</tr>
<tr>
<td>Radium</td>
<td>88</td>
<td>140</td>
<td>88</td>
<td></td>
</tr>
</tbody>
</table>

4.71 a. protons  
   b. protons  
   c. alkali metals
4.73  a. 3, Li  
   b. 86, Rn  
   c. 24.31 amu, Mg  
   d. 19.00 amu, F

4.75  a. Cobalt, Co  
   b. Germanium, Ge  
   c. Sodium, Na  
   d. Cadmium, Cd  
   e. Samarium, Sm

4.77  a. 25 protons, 25 electrons  
   b. 15 protons, 15 electrons  
   c. 38 protons, 38 electrons  
   d. 27 protons, 27 electrons  
   e. 92 protons, 92 electrons

4.79  a. 28 protons, 30 neutrons, 28 electrons  
   b. \(^{63}\)Ni  
   c. \(^{63}\)Cu

4.81  a. \(^1\)H  
   b. \(^{208}\)Hg  
   c. \(^{185}\)Os  
   d. \(^{55}\)Cr

4.83  28.09 amu

4.85  a. 79 protons, 118 neutrons, 79 electrons  
   b. \(^{195}\)Au  
   c. \(^{198}\)Pt

4.87  207.3 amu

4.89  \(8.09 \times 10^7\) atoms of Na
Electronic Structure of Atoms and Periodic Trends

ROBERT AND JENNIFER work in the materials science department of a research laboratory. As materials engineers, they deal with developing and testing various materials that are used in the manufacturing of consumer goods like computer chips, television sets, golf clubs, and snow skis. These engineers create new materials out of metals, ceramics, plastics, semiconductors, and combinations of materials called composites, which are needed for mechanical, chemical, and electrical industries. Often materials engineers study materials at an atomic level to learn how to improve the characteristics of materials.

In their research, Robert and Jennifer are working with some of the elements in Groups 3A (13), 4A (14), and 5A (15) of the periodic table. These elements, such as silicon, have properties that make them good semiconductors. A microchip requires growing a single crystal of a semiconductor such as pure silicon. When small amounts of impurities are added to the crystalline structure, holes form through which electrons can travel with little obstruction. Microchips are manufactured for use in computers, cell phones, satellites, televisions, calculators, GPS, and many other devices. Robert and Jennifer are working on materials that may be used to develop more complex microchips that lead to new applications.

CAREER

Materials Engineer

Materials engineers work with metals, ceramics, plastics, semiconductors, and composites to develop new or improved products such as computer chips, aircraft material, and tennis racquets.

Engineers use mathematics and the principles of chemistry and other sciences to solve technical problems. They may also work in the development and testing of new materials. Materials engineers typically have bachelor’s degrees in mechanical, electrical, or chemical engineering, or related fields.
5.1 Electromagnetic Radiation

**LEARNING GOAL** Compare the wavelength, frequency, and energy of electromagnetic radiation.

When we listen to a radio, use a microwave oven, turn on a light, see the colors of a rainbow, or have an X-ray taken, we are experiencing various forms of electromagnetic radiation. All of these types of electromagnetic radiation, including light, consist of particles that move as waves of energy.

**Wavelength and Frequency**

You are probably familiar with the action of waves in the ocean. If you were at a beach, you might notice that the water in each wave rises and falls as the wave comes in to shore. The highest point on the wave is called a *crest*, whereas the lowest point is a *trough*. On a calm day, there might be long distances between crests or troughs. However, if there is a storm with a lot of energy, the crests or troughs are much closer together.

The waves of electromagnetic radiation also have crests and troughs. The **wavelength** (symbol $\lambda$, lambda) is the distance from a crest or trough in the wave to the next crest or trough in the wave (see **FIGURE 5.1**). In some types of radiation, the crests or troughs are far apart, while in others, they are close together.

![Light source and light passing through a prism](image)

**FIGURE 5.1** (a) Light passing through a prism is separated into a spectrum of colors we see in a rainbow. (b) The wavelength ($\lambda$) is the distance from a crest or trough on a wave to the next crest or trough.

How does the wavelength of red light compare to that of blue light?

The **frequency** (symbol $\nu$, nu) is the number of times the crests of a wave pass a point in 1 s. All electromagnetic radiation travels at the speed of light ($c$), which is a constant value equal to $3.00 \times 10^8$ m/s. Mathematically, the **wave equation** expresses the relationship of the speed of light (m/s) to wavelength (m) and frequency (s$^{-1}$).

$$ c = \lambda \nu \quad \text{Wave equation} $$

Speed of light ($c$) = $3.00 \times 10^8$ m/s = wavelength ($\lambda$) $\times$ frequency ($\nu$)

The speed of light is about a million times faster than the speed of sound, which is the reason we see lightning before we hear thunder during a storm.
Electromagnetic Spectrum

The electromagnetic spectrum is an arrangement of different types of electromagnetic radiation from longest wavelength to shortest wavelength. According to the wave equation, as the wavelength decreases, the frequency increases. Or as the wavelength increases, the frequency decreases. This type of relationship is called an inverse relationship.

Scientists have shown that the energy of electromagnetic radiation is directly related to frequency, which means that energy is also inversely related to the wavelength. Thus, as the wavelength of radiation increases, the frequency and, therefore, the energy decrease.

At one end of the electromagnetic spectrum are radiations with long wavelengths such as radio waves that are used for AM and FM radio bands, cellular phones, and TV signals. The wavelength of a typical AM radio wave can be as long as a football field. Microwaves have shorter wavelengths and higher frequencies than radio waves. Infrared radiation (IR) is responsible for the heat we feel from sunlight and the infrared lamps used to warm food in restaurants. When we change the volume or the station on a TV set, we use a remote control to send infrared impulses to the receiver in the TV. Wireless technology uses radiation with higher frequencies than infrared to connect many electronic devices, including mobile and cell phones and laptops (see FIGURE 5.2).

Visible light with wavelengths from 700 to 400 nm is the only radiation our eyes can detect. Red light has the longest wavelength at 700 nm; orange is about 600 nm; green is about 500 nm; and violet at 400 nm has the shortest wavelength of visible light. We see objects as different colors because the objects reflect only certain wavelengths, which are absorbed by our eyes.

Ultraviolet (UV) light has shorter wavelengths and higher frequencies than violet light of the visible range. The UV radiation in sunlight can cause serious sunburn, which may lead to skin cancer. While some UV light from the Sun is blocked by the ozone layer, the cosmetic industry has developed sunscreens to prevent the absorption of UV light by the skin. X-rays have shorter wavelengths than ultraviolet light, which means they have some of the highest frequencies. X-rays can pass through soft substances but not metals or bone, which allows us to see images of the bones and teeth in the body.

FIGURE 5.2 The electromagnetic spectrum shows the arrangement of wavelengths of electromagnetic radiation. The visible portion consists of wavelengths from 700 nm to 400 nm.

ENGAGE

Which type of electromagnetic radiation has lower frequency, ultraviolet or infrared?
Microwaves with wavelengths of about 1 cm heat water molecules, which heats food.

**SAMPLE PROBLEM 5.1 The Electromagnetic Spectrum**

Arrange the following in order of decreasing wavelengths: X-rays, ultraviolet light, FM radio waves, and microwaves.

**TRY IT FIRST**

**SOLUTION**

The electromagnetic radiation with the longest wavelength is FM radio waves, then microwaves, followed by ultraviolet light, and then X-rays, which have the shortest wavelengths.

**STUDY CHECK 5.1**

Visible light contains colors from red to violet.

a. What color of light has the shortest wavelength?
b. What color of light has the lowest frequency?

**ANSWER**

a. violet light
b. red light

---

**CHEMISTRY LINK TO HEALTH**

**Biological Reactions to UV Light**

Our everyday life depends on sunlight, but exposure to sunlight can have damaging effects on living cells, and too much exposure can even cause their death. Light energy, especially ultraviolet (UV), excites electrons and may lead to unwanted chemical reactions. The list of damaging effects of sunlight includes sunburn; wrinkling; premature aging of the skin; changes in the DNA of the cells, which can lead to skin cancers; inflammation of the eyes; and perhaps cataracts. Some drugs, like the acne medications Accutane and Retin-A, as well as antibiotics, diuretics, sulfonamides, and estrogen, make the skin extremely sensitive to light.

*Phototherapy* uses light to treat certain skin conditions, including psoriasis, eczema, and dermatitis. In the treatment of psoriasis, for example, oral drugs are given to make the skin more photosensitive; then exposure to UV radiation follows. Low-energy radiation (blue light) with wavelengths from 390 to 470 nm is used to treat babies with neonatal jaundice, which converts high levels of bilirubin to water-soluble compounds that can be excreted from the body. Sunlight is also a factor in stimulating the immune system.

In a type of depression called *seasonal affective disorder* or SAD, people experience mood swings and depression during the winter. Some research suggests that SAD is the result of a decrease in serotonin, or an increase in melatonin, when there are fewer hours of sunlight. One treatment for SAD is therapy using bright light provided by a lamp called a light box. A daily exposure to blue light (460 nm) for 30 to 60 min seems to reduce symptoms of SAD.
5.2 Atomic Spectra and Energy Levels

LEARNING GOAL Explain how atomic spectra correlate with the energy levels in atoms.

When the white light from the Sun or a light bulb is passed through a prism or raindrops, it produces a continuous spectrum, like a rainbow. When atoms of elements are heated, they also produce light. At night, you may have seen the yellow color of sodium streetlamps or the red color of neon lights.

Photons
The light emitted from a streetlamp or by atoms that are heated is a stream of particles called photons. Every photon is a packet of energy with both particle and wave characteristics that travels at the speed of light. High-frequency photons have high energy and short wavelengths, whereas low-frequency photons have low energy and long wavelengths.

Photons play an important role in our modern world, particularly in the use of lasers, which use a narrow range of wavelengths. For example, lasers use photons of a single frequency to read pits on compact discs (CDs) and digital versatile discs (DVDs) or to scan bar codes on labels when we buy groceries. A CD is read by a laser with a wavelength of 780 nm. The newer DVDs are read by a blue laser with a wavelength of 405 nm, hence the name Blu-ray. The shorter wavelength allows a smaller pit size on the disc, which means that the disc has a greater storage capacity. In hospitals, high-energy photons are used in treatments to reach tumors within the tissues without damaging the surrounding tissue.
Atomic Spectra

When the light emitted from heated elements is passed through a prism, it does not produce a continuous spectrum. Instead, an atomic spectrum is produced that consists of lines of different colors separated by dark areas (see Figure 5.3). This separation of colors indicates that only certain wavelengths of light are produced when an element is heated, which gives each element a unique atomic spectrum.

![Light passes through a slit](image)

Prism

Film

Strontium light spectrum

Barium light spectrum

**Figure 5.3** In an atomic spectrum, light from a heated element separates into distinct lines.

**Q** Why don't the elements form a continuous spectrum as seen with white light?

Electron Energy Levels

Scientists have now determined that the lines in atomic spectra are associated with changes in the energies of the electrons. In an atom, each electron has a specific energy level, which is assigned a value called the principal quantum number \( n \), \( n = 1, n = 2, \ldots \). Generally, electrons in the lower energy levels are closer to the nucleus, while electrons in the higher energy levels are farther away. The energy of an electron is quantized, which means that the energy of an electron can only have specific energy values, but cannot have values between these values.

**Order of Increasing Energy of Levels**

\[
1 < 2 < 3 < 4 < 5 < 6 < 7
\]

As an analogy, we can think of the energy levels of an atom as similar to the shelves in a bookcase. The first shelf is the lowest energy level; the second shelf is the second energy level; and so on. If we are arranging books on the shelves, it would take less energy to fill the bottom shelf first, and then the second shelf, and so on. However, we could never get any book to stay in the space between any of the shelves. Similarly, the energy of an electron must be at a specific energy level, and not between.

Unlike standard bookcases, however, there is a large difference between the energy of the first and second energy levels, but then the higher energy levels are closer together. Another difference is that the lower electron energy levels hold fewer electrons than the higher energy levels.
Changes in Energy Levels

An electron can change from one energy level to a higher energy level only if it absorbs the energy equal to the difference in energy levels. When an electron changes to a lower energy level, it emits energy equal to the difference between the two energy levels. If the energy emitted is in the visible range, we see one of the colors of visible light (see Figure 5.4). The yellow color of sodium streetlights and the red color of neon lights are examples of electrons emitting energy in the visible color range.

Colors are produced when electricity excites electrons in noble gases.

SAMPLE PROBLEM 5.2 Change in Energy Levels

a. How does an electron move to a higher energy level?
b. When an electron drops to a lower energy level, how is energy lost?

TRY IT FIRST

SOLUTION

a. An electron moves to a higher energy level when it absorbs an amount of energy equal to the difference in energy levels.
b. Energy, equal to the difference in energy levels, is emitted as a photon when an electron drops to a lower energy level.

STUDY CHECK 5.2

Why did scientists propose that electrons occupy specific energy levels in an atom?

ANSWER

Because the spectra of elements consisted of only discrete, separated lines, scientists concluded that electrons occupied only certain energy levels in the atom.

CHEMISTRY LINK TO THE ENVIRONMENT

Energy-Saving Fluorescent Bulbs

Compact fluorescent lights (CFL) are replacing the standard light bulb we use in our homes and workplaces. Compared to a standard light bulb, the CFL has a longer life and uses less electricity. Within about 20 days of use, the fluorescent bulb saves enough money in electricity to pay for its higher initial cost.

A standard incandescent light bulb has a thin tungsten filament inside a sealed glass bulb. When the light is switched on, electricity flows through this filament, and electrical energy is converted to heat energy. When the filament reaches a temperature around 2300 °C, we see white light.
A fluorescent bulb produces light in a different way. When the switch is turned on, electrons move between two electrodes and collide with mercury atoms in a mixture of mercury and argon gas inside the bulb. When the electrons in the mercury atoms absorb energy from the collisions, they are raised to higher energy levels. As electrons fall to lower energy levels emitting ultraviolet radiation, they strike the phosphor coating inside the tube, and fluorescence occurs as visible light is emitted.

The production of light is more efficient in a fluorescent bulb than in an incandescent light bulb. A 75-watt incandescent bulb can be replaced by a 20-watt CFL that gives the same amount of light, providing a 70% reduction in electricity costs. A typical incandescent light bulb lasts one to two months, whereas a compact fluorescent light bulb lasts from 1 to 2 yr.

One drawback of the CFL is that each contains about 4 mg of mercury. As long as the bulb stays intact, no mercury is released. However, used CFL bulbs should not be disposed of in household trash but rather should be taken to a recycling center.

5.18 Electrons drop to lower energy levels when they _____ (absorb/emit) a photon.

5.19 Identify the photon in each pair with the greater energy.
   a. green light or yellow light
   b. red light or blue light

5.20 Identify the photon in each pair with the greater energy.
   a. orange light or violet light
   b. infrared light or ultraviolet light

5.3 Sublevels and Orbitals

LEARNING GOAL Describe the sublevels and orbitals for the electrons in an atom.

We have seen that the protons and neutrons are contained in the small, dense nucleus of an atom. However, it is the electrons within the atoms that determine the physical and chemical properties of the elements. Therefore, we will look at the arrangement of electrons within the large volume of space surrounding the nucleus.

There is a limit to the number of electrons allowed in each energy level. Only a few electrons can occupy the lower energy levels, while more electrons can be accommodated in higher energy levels. The maximum number of electrons allowed in any energy level is calculated using the formula \(2n^2\) (two times the square of the principal quantum number). TABLE 5.1 shows the maximum number of electrons allowed in the first four energy levels.

**TABLE 5.1 Maximum Number of Electrons Allowed in Energy Levels 1 to 4**

<table>
<thead>
<tr>
<th>Energy Level ((n))</th>
<th>1</th>
<th>2((n^2))</th>
<th>3((n^2))</th>
<th>4((n^2))</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>2(1)^2</td>
<td>2(2)^2</td>
<td>2(3)^2</td>
</tr>
<tr>
<td>Maximum Number of Electrons</td>
<td>2</td>
<td>8</td>
<td>18</td>
<td>32</td>
</tr>
</tbody>
</table>

**Sublevels**

Each of the energy levels consists of one or more sublevels, in which electrons with identical energy are found. The sublevels are identified by the letters \(s, p, d,\) and \(f\). The number of sublevels within an energy level is equal to the principal quantum number, \(n\). For example, the first energy level \((n = 1)\) has only one sublevel, \(1s\). The second energy level \((n = 2)\) has two sublevels, \(2s\) and \(2p\). The third energy level \((n = 3)\) has three sublevels, \(3s, 3p,\) and \(3d\). The fourth energy level \((n = 4)\) has four sublevels, \(4s, 4p, 4d,\) and \(4f\). Energy levels \(n = 5, n = 6,\) and \(n = 7\) also have as many sublevels as the value of \(n\), but only \(s, p, d,\) and \(f\) sublevels are needed to hold the electrons in atoms of the 118 known elements (see FIGURE 5.5).
5.3 Sublevels and Orbitals

Within each energy level, the $s$ sublevel has the lowest energy. If there are additional sublevels, the $p$ sublevel has the next lowest energy, then the $d$ sublevel, and finally the $f$ sublevel.

**Order of Increasing Energy of Sublevels in an Energy Level**

$s < p < d < f$

<table>
<thead>
<tr>
<th>Energy Level</th>
<th>Number of Sublevels</th>
<th>Types of Sublevels</th>
</tr>
</thead>
<tbody>
<tr>
<td>$n = 4$</td>
<td>4</td>
<td>$s$</td>
</tr>
<tr>
<td>$n = 3$</td>
<td>3</td>
<td>$s$</td>
</tr>
<tr>
<td>$n = 2$</td>
<td>2</td>
<td>$s$</td>
</tr>
<tr>
<td>$n = 1$</td>
<td>1</td>
<td>$s$</td>
</tr>
</tbody>
</table>

Orbitals

There is no way to know the exact location of an electron in an atom. Instead, scientists describe the location of an electron in terms of probability. The **orbital** is the three-dimensional volume in which electrons have the highest probability of being found.

As an analogy, imagine that you could draw a circle with a 100-m radius around your chemistry classroom. There is a high probability of finding you within that circle when your chemistry class is in session. But once in a while, you may be outside that circle because you were sick or your car did not start.

**Shapes of Orbitals**

Each type of orbital has a unique three-dimensional shape. Electrons in an $s$ orbital are most likely found in a region with a spherical shape. Imagine that you take a picture of the location of an electron in an $s$ orbital every second for an hour. When all these pictures are overlaid, the result, called a **probability density**, would look like the electron cloud shown in **Figure 5.5a**. For convenience, we draw this electron cloud as a sphere called an $s$ orbital. There is one $s$ orbital for every energy level starting with $n = 1$. For example, in the first, second, and third energy levels there are $s$ orbitals designated as $1s$, $2s$, and $3s$. As the principal quantum number increases, there is an increase in the size of the $s$ orbitals, although the shape is the same (see **Figure 5.5b**).

The orbitals occupied by $p$, $d$, and $f$ electrons have three-dimensional shapes different from those of the $s$ electrons. There are three $p$ orbitals, starting with $n = 2$. Each $p$ orbital has two lobes like a balloon tied in the middle. The three $p$ orbitals are arranged in three perpendicular directions, along the $x$, $y$, and $z$ axes around the nucleus (see **Figure 5.6**). As with $s$ orbitals, the shape of $p$ orbitals is the same, but the volume increases at higher energy levels.

In summary, the $n = 2$ energy level, which has $2s$ and $2p$ sublevels, consists of one $s$ orbital and three $p$ orbitals.

**Figure 5.5** ▶ The number of sublevels in an energy level is the same as the principal quantum number, $n$.

**Question:** How many sublevels are in the $n = 3$ energy level?

**Figure 5.6** ▶ (a) The electron cloud of an $s$ orbital represents the highest probability of finding an $s$ electron. (b) The $s$ orbitals are shown as spheres. The sizes of the $s$ orbitals increase because they contain electrons at higher energy levels.

**Question:** Is the probability high or low of finding an $s$ electron outside an $s$ orbital?
What are some similarities and differences of the \( p \) orbitals in the \( n = 3 \) energy level?

Energy level \( n = 3 \) consists of three sublevels \( s \), \( p \), and \( d \). The \( d \) sublevels contain five \( d \) orbitals (see Figure 5.8).

How many orbitals are there in the \( 5d \) sublevel?

Energy level \( n = 4 \) consists of four sublevels \( s \), \( p \), \( d \), and \( f \). In the \( f \) sublevel, there are seven \( f \) orbitals. The shapes of \( f \) orbitals are complex and so we have not included them in this text.

**SAMPLE PROBLEM 5.3 Energy Levels, Sublevels, and Orbitals**

Indicate the type and number of orbitals in each of the following energy levels or sublevels:

a. \( 3p \) sublevel  

b. \( n = 2 \)  

c. \( n = 3 \)  

d. \( 4d \) sublevel  

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>energy level, sublevel</td>
<td>orbitals</td>
<td>( n ), orbitals in ( s, p, d, f )</td>
<td></td>
</tr>
</tbody>
</table>
a. The $3p$ sublevel contains three $3p$ orbitals.
b. The $n = 2$ energy level consists of one $2s$ and three $2p$ orbitals.
c. The $n = 3$ energy level consists of one $3s$, three $3p$, and five $3d$ orbitals.
d. The $4d$ sublevel contains five $4d$ orbitals.

**STUDY CHECK 5.3**

What is similar and what is different for $1s$, $2s$, and $3s$ orbitals?

**ANSWER**

The $1s$, $2s$, and $3s$ orbitals are all spherical, but they increase in volume because the electron is most likely to be found farther from the nucleus for higher energy levels.

**Orbital Capacity and Electron Spin**

The *Pauli exclusion principle* states that each orbital can hold a maximum of two electrons. According to a model for electron behavior, an electron is seen as spinning on its axis, which generates a magnetic field. When two electrons are in the same orbital, they will repel each other unless their magnetic fields cancel. This happens only when the two electrons spin in opposite directions. We can represent the spins of the electrons in the same orbital with one arrow pointing up and the other pointing down.

![Electron spin arrows](image)

**Number of Electrons in Sublevels**

There is a maximum number of electrons that can occupy each sublevel. An $s$ sublevel holds one or two electrons. Because each $p$ orbital can hold up to two electrons, the three $p$ orbitals in a $p$ sublevel can accommodate six electrons. A $d$ sublevel with five $d$ orbitals can hold a maximum of 10 electrons. With seven $f$ orbitals, an $f$ sublevel can hold up to 14 electrons.

As mentioned earlier, higher energy levels such as $n = 5$, $6$, and $7$ would have 5, 6, and 7 sublevels, but those beyond sublevel $f$ are not utilized by the atoms of the elements known today. The total number of electrons in all the sublevels adds up to give the electrons allowed in an energy level. The number of sublevels, the number of orbitals, and the maximum number of electrons for energy levels 1 to 4 are shown in [*TABLE 5.2*](#).

**Table 5.2** Electron Capacity in Sublevels for Energy Levels 1 to 4

<table>
<thead>
<tr>
<th>Energy Level (n)</th>
<th>Number of Sublevels</th>
<th>Type of Sublevel</th>
<th>Number of Orbitals</th>
<th>Maximum Number of Electrons</th>
<th>Total Electrons $(2n^2)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>4</td>
<td>4</td>
<td>$4f$</td>
<td>7</td>
<td>14</td>
<td>32</td>
</tr>
<tr>
<td></td>
<td></td>
<td>$4d$</td>
<td>5</td>
<td>10</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>$4p$</td>
<td>3</td>
<td>6</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>$4s$</td>
<td>1</td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>$3d$</td>
<td>5</td>
<td>10</td>
<td>18</td>
</tr>
<tr>
<td></td>
<td></td>
<td>$3p$</td>
<td>3</td>
<td>6</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>$3s$</td>
<td>1</td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>$2p$</td>
<td>3</td>
<td>6</td>
<td>8</td>
</tr>
<tr>
<td></td>
<td></td>
<td>$2s$</td>
<td>1</td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>1</td>
<td>1</td>
<td>$1s$</td>
<td>1</td>
<td>2</td>
<td>2</td>
</tr>
</tbody>
</table>
Chapter 5: Electronic Structure of Atoms and Periodic Trends

5.4 Orbital Diagrams and Electron Configurations

**LEARNING GOAL** Draw the orbital diagram and write the electron configuration for an element.

We can now look at how electrons are arranged within an atom. Electrons are added first to orbitals with the lowest energy levels, building progressively by adding electrons to levels with higher energies. This process of building the electrons in an atom is known as the **Aufbau principle**.

In an **orbital diagram**, electrons are shown as arrows that are placed in boxes that represent the orbitals in order of increasing energy (see **Figure 5.9**).

To draw an orbital diagram, the lowest energy orbitals are filled first. For example, we can draw the orbital diagram for carbon. The atomic number of carbon is 6, which means that a carbon atom has six electrons. The first two electrons go into the 1s orbital; the next two electrons go into the 2s orbital. In the orbital diagram, the two electrons in the 1s and 2s orbitals are shown with opposite spins; the first arrow is up and the second is down. The last two electrons in carbon begin to fill the 2p sublevel, which has the next lowest energy. However, there are three 2p orbitals of equal energy. Because the negatively charged electrons repel each other, they are placed in separate 2p orbitals. **Hund’s rule** states that there is less repulsion when electrons are placed in separate orbitals of the same sublevel. With few exceptions (which will be noted later in this chapter) lower energy sublevels are filled first, and then the “building” of electrons continues to the next lowest energy sublevel that is available until all the electrons are placed.

When do we need to use Hund’s rule in drawing orbital diagrams?

<table>
<thead>
<tr>
<th>Orbital</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>1s</td>
<td>1s 1s</td>
</tr>
<tr>
<td>2s</td>
<td>2s 2s</td>
</tr>
<tr>
<td>2p</td>
<td>2p 2p 2p 2p</td>
</tr>
</tbody>
</table>

**Orbital diagram for carbon**

**Halffilled**  **Empty**
Electron Configurations

Chemists use a notation called the **electron configuration** to indicate the placement of the electrons of an atom in order of increasing energy. As in the orbital diagrams, an electron configuration is written with the lowest energy sublevel first, followed by the next lowest energy sublevel. The number of electrons in each sublevel is shown as a superscript.

**Electron Configuration for Carbon**

<table>
<thead>
<tr>
<th>Type of orbital</th>
<th>Number of electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>1s</td>
<td>1s^2 2s^2 2p^2</td>
</tr>
</tbody>
</table>

Read as “one s two, two s two, two p two”

**Period 1: Hydrogen and Helium**

We will draw orbital diagrams and write the corresponding electron configurations for the elements H and He in Period 1. The 1s orbital (which is the 1s sublevel) is written first because it has the lowest energy. Hydrogen has one electron in the 1s sublevel; helium has two. In the orbital diagram, the electrons for helium are shown as arrows pointing in opposite directions.
ChApter 5
Electronic Structure of Atoms and Periodic Trends

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Number</th>
<th>Orbital Diagram</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1</td>
<td>1s</td>
<td>1s^1</td>
</tr>
<tr>
<td>He</td>
<td>2</td>
<td>1s</td>
<td>1s^2</td>
</tr>
</tbody>
</table>

Period 2: Lithium to Neon

Period 2 begins with lithium, which has three electrons. The first two electrons fill the 1s orbital, whereas the third electron goes into the 2s orbital, the sublevel with the next lowest energy. In beryllium, another electron is added to complete the 2s orbital. The next six electrons are used to fill the 2p orbitals. The electrons are added to separate p orbitals (Hund’s rule) from boron to nitrogen, which gives three half-filled 2p orbitals.

From oxygen to neon, the remaining three electrons are paired up using opposite spins to complete the 2p sublevel. In writing the electron configurations for the elements in Period 2, begin with the 1s orbital followed by the 2s and then the 2p orbitals.

An electron configuration can also be written in an abbreviated configuration. The electron configuration of the preceding noble gas is replaced by writing its element symbol inside square brackets. For example, the electron configuration for lithium, 1s^22s^1, can be abbreviated as [He]2s^1 where [He] replaces 1s^2.

---

**Guide to Drawing Orbital Diagrams**

**STEP 1**
Draw boxes to represent the occupied orbitals.

**STEP 2**
Place a pair of electrons with opposite spins in each filled orbital.

**STEP 3**
Place the remaining electrons in the last occupied sublevel in separate orbitals.

**SAMPLE PROBLEM 5.4  Drawing Orbital Diagrams**

Nitrogen is an element that is used in the formation of amino acids, proteins, and nucleic acids. Draw the orbital diagram for nitrogen.

**TRY IT FIRST**

**SOLUTION**

**STEP 1**
Draw boxes to represent the occupied orbitals. Nitrogen has atomic number 7, which means it has seven electrons. For the orbital diagram, we draw boxes to represent the 1s, 2s, and 2p orbitals.

```
1s 2s 2p
```

---
**STEP 2** Place a pair of electrons with opposite spins in each filled orbital. First, we place a pair of electrons with opposite spins in both the 1s and 2s orbitals.

```
1s  2s  2p
↑   ↓   □□□
```

**STEP 3** Place the remaining electrons in the last occupied sublevel in separate orbitals. Then we place the three remaining electrons in three separate 2p orbitals with arrows drawn in the same direction.

```
1s  2s  2p
↑   ↑   ▲▲▲
```

Orbital diagram for nitrogen (N)

**STUDY CHECK 5.4**
Draw the orbital diagram for fluorine, which is used to make nonstick coatings for cookware.

**ANSWER**
```
1s  2s  2p
↑   ↑   ↑↑↑
```

**Period 3: Sodium to Argon**
In Period 3, electrons enter the orbitals of the 3s and 3p sublevels, but not the 3d sublevel. We notice that the elements sodium to argon, which are directly below the elements lithium to neon in Period 2, have a similar pattern of filling their s and p orbitals. In sodium and magnesium, one and two electrons go into the 3s orbital. The electrons for aluminum, silicon, and phosphorus go into separate 3p orbitals. The remaining electrons in sulfur, chlorine, and argon are paired up (opposite spins) with the electrons already in the 3p orbitals. For the abbreviated electron configurations of Period 3, the symbol [Ne] replaces 1s\(^2\)2s\(^2\)2p\(^6\).

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Number</th>
<th>Orbital Diagram (3s and 3p orbitals only)</th>
<th>Electron Configuration</th>
<th>Abbreviated Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>11</td>
<td>[Ne] 3(^s) 3(^p)</td>
<td>1s(^2)2s(^2)2p(^6)3s(^1)</td>
<td>[Ne]3s(^1)</td>
</tr>
<tr>
<td>Mg</td>
<td>12</td>
<td>[Ne] 3(^s) 3(^p)</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)</td>
<td>[Ne]3s(^2)</td>
</tr>
<tr>
<td>Al</td>
<td>13</td>
<td>[Ne] 3(^s) 3(^p)</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^1)</td>
<td>[Ne]3s(^2)3p(^1)</td>
</tr>
<tr>
<td>Si</td>
<td>14</td>
<td>[Ne] 3(^s) 3(^p)</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^2)</td>
<td>[Ne]3s(^2)3p(^2)</td>
</tr>
<tr>
<td>P</td>
<td>15</td>
<td>[Ne] 3(^s) 3(^p)</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^3)</td>
<td>[Ne]3s(^2)3p(^3)</td>
</tr>
<tr>
<td>S</td>
<td>16</td>
<td>[Ne] 3(^s) 3(^p)</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^4)</td>
<td>[Ne]3s(^2)3p(^4)</td>
</tr>
<tr>
<td>Cl</td>
<td>17</td>
<td>[Ne] 3(^s) 3(^p)</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^5)</td>
<td>[Ne]3s(^2)3p(^5)</td>
</tr>
<tr>
<td>Ar</td>
<td>18</td>
<td>[Ne] 3(^s) 3(^p)</td>
<td>1s(^2)2s(^2)2p(^6)3s(^2)3p(^6)</td>
<td>[Ne]3s(^2)3p(^6)</td>
</tr>
</tbody>
</table>
SAMPLE PROBLEM 5.5  Electron Configurations

Silicon is the basis of semiconductors. Write the complete and abbreviated electron configurations for silicon.

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>silicon (Si)</td>
<td>electron configuration,</td>
<td>periodic table,</td>
</tr>
<tr>
<td></td>
<td>abbreviated electron</td>
<td>atomic number,</td>
</tr>
<tr>
<td></td>
<td>configuration</td>
<td>order of filling orbitals</td>
</tr>
</tbody>
</table>

STEP 1  State the number of electrons from the atomic number on the periodic table. Silicon has an atomic number of 14, which means it has 14 electrons.

STEP 2  Write the number of electrons for each orbital in order of increasing energy until filling is complete.

\[1s^2 2s^2 2p^6 3s^2 3p^2\]  Electron configuration for Si

STEP 3  Write an abbreviated electron configuration by replacing the configuration of the preceding noble gas with its symbol.

\[[\text{Ne}]3s^2 3p^2\]  Abbreviated electron configuration for Si

STUDY CHECK 5.5

Write the complete and abbreviated electron configurations for sulfur, which is a macro-mineral in proteins, vitamin B1, and insulin.

ANSWER

\[1s^2 2s^2 2p^6 3s^2 3p^4, [\text{Ne}]3s^2 3p^4\]

QUESTIONS AND PROBLEMS

5.4 Orbital Diagrams and Electron Configurations

LEARNING GOAL  Draw the orbital diagram and write the electron configuration for an element.

5.29  Compare the terms electron configuration and abbreviated electron configuration.

5.30  Compare the terms orbital diagram and electron configuration.

5.31  Draw the orbital diagram for each of the following:

a. boron  b. aluminum  c. phosphorus  d. argon

5.32  Draw the orbital diagram for each of the following:

a. carbon  b. sulfur  c. magnesium  d. beryllium

5.33  Write the complete electron configuration for each of the following:

a. nickel  b. sodium  c. lithium  d. titanium

5.34  Write the complete electron configuration for each of the following:

a. nitrogen  b. chlorine  c. strontium  d. neon

5.35  Write the abbreviated electron configuration for each of the following:

a. tin  b. cadmium  c. selenium  d. fluorine

5.36  Write the abbreviated electron configuration for each of the following:

a. barium  b. oxygen  c. manganese  d. arsenic

5.37  Give the symbol of the element with each of the following electron or abbreviated electron configurations:

a. \(1s^2 2s^1\)  b. \(1s^2 2s^2 2p^6 3s^1 3p^6 4s^2 3d^2\)  c. \([\text{Ar}]4s^2 3d^{10} 4p^2\)  d. \([\text{He}]2s^2 2p^6\)

5.38  Give the symbol of the element with each of the following electron or abbreviated electron configurations:

a. \(1s^2 2s^2 2p^4\)  b. \([\text{Ne}]3s^2\)  c. \(1s^2 2s^2 2p^6 3s^2 3p^6\)  d. \([\text{Ne}]3s^2 3p^1\)

5.39  Give the symbol of the element that meets the following conditions:

a. has three electrons in the \(n = 3\) energy level  b. has two \(2p\) electrons  c. completes the \(3p\) sublevel  d. completes the \(2s\) sublevel

5.40  Give the symbol of the element that meets the following conditions:

a. has five electrons in the \(3p\) sublevel  b. has four \(2p\) electrons  c. completes the \(3s\) sublevel  d. has one electron in the \(3s\) sublevel
5.5 Electron Configurations and the Periodic Table

**LEARNING GOAL** Write the electron configuration for an atom using the sublevel blocks on the periodic table.

Up to now, we have written electron configurations using their energy level diagrams. As configurations involve more energy levels, this can become tedious. However, the electron configurations of the elements are related to their position on the periodic table. Different sections or blocks within the periodic table correspond to the s, p, d, and f sublevels (see Figure 5.10). Therefore, we can “build” the electron configurations of atoms by reading the periodic table in order of increasing atomic number.

### Blocks on the Periodic Table

1. The **s block** includes hydrogen and helium as well as the elements in Group 1A (1) and Group 2A (2). This means that the final one or two electrons in the elements of the s block are located in an s orbital. The period number indicates the particular s orbital that is filling: 1s, 2s, and so on.

2. The **p block** consists of the elements in Group 3A (13) to Group 8A (18). There are six p block elements in each period because three p orbitals can hold up to six electrons. The period number indicates the particular p sublevel that is filling: 2p, 3p, and so on.

3. The **d block**, containing the transition elements, first appears after calcium (atomic number 20). There are 10 elements in each period of the d block because five d orbitals can hold up to 10 electrons. The particular d sublevel is one less \((n-1)\) than the period number. For example, in Period 4, the d block is the 3d sublevel. In Period 5, the d block is the 4d sublevel.

4. The **f block**, the inner transition elements, are the two rows at the bottom of the periodic table. There are 14 elements in each f block because seven f orbitals can hold up to 14 electrons. Elements that have atomic numbers higher than 57 (La) have electrons in the 4f block. The particular f sublevel is two less \((n-2)\) than the period number. For example, in Period 6, the f block is the 4f sublevel. In Period 7, the f block is the 5f sublevel.

### Writing Electron Configurations Using Sublevel Blocks

Now we can write electron configurations using the sublevel blocks on the periodic table. As before, each configuration begins at H. But now we move across the table from left to right, writing down each sublevel block we come to until we reach the element for which

---

**Figure 5.10** Electron configurations follow the order of occupied sublevels on the periodic table.

- If neon is in Group 8A (18), Period 2, how many electrons are in the 1s, 2s, and 2p sublevels of neon?

---

**ENGAGE**

On the periodic table, what are the group numbers that make up the p block? the s block?

**CORE CHEMISTRY SKILL**

Using the Periodic Table to Write Electron Configurations
we are writing an electron configuration. We will show how to write the electron configuration for chlorine (atomic number 17) from the sublevel blocks on the periodic table in Sample Problem 5.6.

**SAMPLE PROBLEM 5.6 Using the Sublevel Blocks to Write Electron Configurations**

Use the sublevel blocks on the periodic table to write the electron configuration for chlorine.

**TRY IT FIRST**

**SOLUTION**

**STEP 1** Locate the element on the periodic table. Chlorine (atomic number 17) is in Group 7A (17) and Period 3.

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>element Cl</td>
<td>electron configuration</td>
<td>sublevel blocks</td>
</tr>
</tbody>
</table>

**STEP 2** Write the filled sublevels in order, going across each period.

<table>
<thead>
<tr>
<th>Period</th>
<th>Sublevel Block Filling</th>
<th>Sublevel Block Notation (filled)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1s (H → He)</td>
<td>1s²</td>
</tr>
<tr>
<td>2</td>
<td>2s (Li → Be) and 2p (B → Ne)</td>
<td>2s²2p⁶</td>
</tr>
<tr>
<td>3</td>
<td>3s (Na → Mg)</td>
<td>3s²</td>
</tr>
</tbody>
</table>

**STEP 3** Complete the configuration by counting the electrons in the last occupied sublevel block. Because chlorine is the fifth element in the 3p block, there are five electrons in the 3p sublevel.

<table>
<thead>
<tr>
<th>Period</th>
<th>Sublevel Block Filling</th>
<th>Sublevel Block Notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>3p (Al → Cl)</td>
<td>3p⁵</td>
</tr>
</tbody>
</table>

The electron configuration for chlorine (Cl) is: 1s²2s²2p⁶3s²3p⁵.

**STUDY CHECK 5.6**

Use the sublevel blocks on the periodic table to write the electron configuration for argon.

**ANSWER**

1s²2s²2p⁶3s²3p⁶

**Electron Configurations for Period 4 and Above**

Up to Period 4, the filling of the sublevels has progressed in order. However, if we look at the sublevel blocks in Period 4, we see that the 4s sublevel fills before the 3d sublevel. This occurs because the electrons in the 4s sublevel have slightly lower energy than the electrons in the 3d sublevel. This order occurs again in Period 5 when the 5s sublevel fills before the 4d sublevel and again in Period 6 when the 6s fills before the 5d.

At the beginning of Period 4, the electrons in potassium (19) and calcium (20) go into the 4s sublevel. In scandium, the next electron added after the 4s sublevel is filled goes into the 3d block. The 3d block continues to fill until it is complete with 10 electrons at zinc (30). Once the 3d block is complete, the next six electrons, gallium to krypton, go into the 4p block.
### Sample Problem 5.7 Using Sublevel Blocks to Write Electron Configurations

Selenium is used in making glass and in pigments. Use the sublevel blocks on the periodic table to write the electron configuration for selenium.

**Try it First**

**Analyze the Problem**

Given:
- Selenium

Need:
- Electron configuration

Connect:
- Sublevel blocks

**Solution**

1. **Step 1** Locate the element on the periodic table. Selenium is in Group 6A (16) and Period 4.
**STEP 2** Write the filled sublevels in order, going across each period.

<table>
<thead>
<tr>
<th>Period</th>
<th>Sublevel Block Filling</th>
<th>Sublevel Block Notation (filled)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1s (H (\rightarrow) He)</td>
<td>1(s^2)</td>
</tr>
<tr>
<td>2</td>
<td>2s (Li (\rightarrow) Be) and 2p (B (\rightarrow) Ne)</td>
<td>2(s^22p^6)</td>
</tr>
<tr>
<td>3</td>
<td>3s (Na (\rightarrow) Mg) and 3p (Al (\rightarrow) Ar)</td>
<td>3(s^23p^6)</td>
</tr>
<tr>
<td>4</td>
<td>4s (K (\rightarrow) Ca) and 3d (Sc (\rightarrow) Zn)</td>
<td>4(s^23d^{10})</td>
</tr>
</tbody>
</table>

**STEP 3** Complete the configuration by counting the electrons in the last occupied sublevel block. Because selenium is the fourth element in the 4p block, there are four electrons in the 4p sublevel.

<table>
<thead>
<tr>
<th>Period</th>
<th>Sublevel Block Filling</th>
<th>Sublevel Block Notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>4</td>
<td>4p (Ga (\rightarrow) Se)</td>
<td>4(p^4)</td>
</tr>
</tbody>
</table>

The electron configuration for selenium (Se) is: \(1s^22s^22p^63s^23p^64s^23d^{10}4p^4\).

**STUDY CHECK 5.7**
Iodine is a micromineral needed for thyroid function. Use the sublevel blocks on the periodic table to write the electron configuration for iodine.

**ANSWER**
\(1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^5\)

**Some Exceptions in Sublevel Block Order**
Within the filling of the 3d sublevel, exceptions occur for chromium and copper. In Cr and Cu, the 3d sublevel is close to being a half-filled or filled sublevel, which is particularly stable. Thus, the electron configuration for chromium has only one electron in the 4s and five electrons in the 3d sublevel to give the added stability of a half-filled d sublevel. This is shown in the abbreviated orbital diagram for chromium:

```
[Ar]  4s      3d (half-filled)      4p
```

A similar exception occurs when copper achieves a stable, filled 3d sublevel with 10 electrons and only one electron in the 4s orbital. This is shown in the abbreviated orbital diagram for copper:

```
[Ar]  4s      3d (filled)        4p
```

After the 4s and 3d sublevels are completed, the 4p sublevel fills as expected from gallium to krypton, the noble gas that completes Period 4. There are also exceptions in filling for the higher d and f electron sublevels, some caused by the added stability of half-filled shells and others where the cause is not known.
5.6 Trends in Periodic Properties

LEARNING GOAL Use the electron configurations of elements to explain the trends in periodic properties.

The electron configurations of atoms are an important factor in the physical and chemical properties of the elements and in the properties of the compounds that they form. In this section, we will look at the valence electrons in atoms, the trends in atomic size, ionization energy, and metallic character. Going across a period, there is a pattern of regular change in these properties from one group to the next. Known as periodic properties, each property increases or decreases across a period, and then the trend is repeated in each successive period. We can use the seasonal changes in temperatures as an analogy for periodic properties. In the winter, temperatures are cold and become warmer in the spring. By summer, the outdoor temperatures are hot but begin to cool in the fall. By winter, we expect cold temperatures again as the pattern of decreasing and increasing temperatures repeats for another year.

Group Number and Valence Electrons

The chemical properties of representative elements are mostly due to the valence electrons, which are the electrons in the outermost energy level. These valence electrons occupy the s and p sublevels with the highest principal quantum number n. The group numbers indicate the number of valence (outer) electrons for the elements in each vertical column. For example, the elements in Group 1A (1), such as lithium, sodium, and...
SOLVING EQUATIONS

KEY MATH SKILLS

IDENTIFYING FUNCTIONAL GROUPS

TOXICOLOGY AND RISK—BENEFIT ASSESSMENT

Try It First

GREENHOUSE GASES

Chemistry Link to

TOXICOLOGY AND RISK—BENEFIT ASSESSMENT

ENGAGE

Engage

Health

b. 4A (14) 6A (16) ÷ 5A (15) - 7A (17) 3A (13) c.

Environment

What are the group numbers, the periods, the number of valence electrons, and the valence electron configurations for sulfur and strontium?

ANSWER

Sulfur is in Group 6A (16), Period 3, has six valence electrons, and a 3s²3p⁴ valence electron configuration. Strontium is in Group 2A (2), Period 5, has two valence electrons, and a 5s² valence electron configuration.

Lewis Symbols

A Lewis symbol is a convenient way to represent the valence electrons, which are shown as dots placed on the sides, top, or bottom of the symbol for the element. One to four valence electrons are arranged as single dots. When there are five to eight electrons, one or
more electrons are paired. Any of the following would be an acceptable Lewis symbol for magnesium, which has two valence electrons:

**Lewis Symbols for Magnesium**

\[ \cdot\text{Mg} \cdot\cdot\text{Mg} \cdot\cdot\cdot\text{Mg} \] \[ \cdot\text{Mg} \cdot\cdot\cdot\text{Mg} \]

Lewis symbols for selected elements are given in **TABLE 5.4**.

---

**TABLE 5.4** Lewis Symbols for Selected Elements in Periods 1 to 4

<table>
<thead>
<tr>
<th>Group Number</th>
<th>1A (1)</th>
<th>2A (2)</th>
<th>3A (13)</th>
<th>4A (14)</th>
<th>5A (15)</th>
<th>6A (16)</th>
<th>7A (17)</th>
<th>8A (18)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of Valence Electrons</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>5</td>
<td>6</td>
<td>7</td>
<td>8*</td>
</tr>
<tr>
<td>Lewis Symbol</td>
<td>H(\cdot)</td>
<td>He(\cdot)</td>
<td>Li(\cdot)</td>
<td>Be(\cdot)</td>
<td>B(\cdot)</td>
<td>C(\cdot\cdot\cdot)</td>
<td>N(\cdot\cdot)</td>
<td>O(\cdot\cdot\cdot)</td>
</tr>
<tr>
<td></td>
<td>Na(\cdot\cdot\cdot)</td>
<td>Mg(\cdot\cdot\cdot)</td>
<td>Al(\cdot\cdot\cdot)</td>
<td>Si(\cdot\cdot\cdot)</td>
<td>P(\cdot\cdot\cdot)</td>
<td>S(\cdot\cdot\cdot)</td>
<td>Cl(\cdot\cdot\cdot)</td>
<td>Ar(\cdot\cdot\cdot)</td>
</tr>
<tr>
<td></td>
<td>K(\cdot\cdot\cdot)</td>
<td>Ca(\cdot\cdot\cdot)</td>
<td>Ga(\cdot\cdot\cdot)</td>
<td>Ge(\cdot\cdot\cdot)</td>
<td>As(\cdot\cdot\cdot)</td>
<td>Se(\cdot\cdot\cdot)</td>
<td>Br(\cdot\cdot\cdot)</td>
<td>Kr(\cdot\cdot\cdot)</td>
</tr>
</tbody>
</table>

*Helium (He) is stable with two valence electrons.

---

**SAMPLE PROBLEM 5.9 Drawing Lewis Symbols**

Draw the Lewis symbol for each of the following:

a. bromine 

b. aluminum

**TRY IT FIRST**

**SOLUTION**

a. The Lewis symbol for bromine, which is in Group 7A (17), has seven valence electrons. Thus, three pairs of dots and one single dot are drawn on the sides of the Br symbol.

\[ \cdot\text{Br} \cdot\cdot\cdot\text{Br} \]

b. The Lewis symbol for aluminum, which is in Group 3A (13), has three valence electrons drawn as single dots on the sides of the Al symbol.

\[ \cdot\text{Al} \cdot\cdot\cdot\text{Al} \]

**STUDY CHECK 5.9**

Draw the Lewis symbol for phosphorus, a macromineral needed for bones and teeth.

**ANSWER**

\[ \cdot\text{P} \cdot\cdot\cdot\text{P} \]

---

**Atomic Size**

The **atomic size** of an atom is determined by the distance of the valence electrons from the nucleus. For each group of representative elements, the atomic size increases going from the top to the bottom because the outermost electrons in each energy level are farther from the nucleus. For example, in Group 1A (1), Li has a valence electron in energy level 2; Na has a valence electron in energy level 3; and K has a valence electron in energy level 4. This means that a K atom is larger than a Na atom and a Na atom is larger than a Li atom (see **FIGURE 5.11**).
The atomic size of representative elements is affected by the attractive forces of the protons in the nucleus on the electrons in the outermost level. For the elements going across a period, the increase in the number of protons in the nucleus increases the positive charge of the nucleus. As a result, the electrons are pulled closer to the nucleus, which means that the atomic size of representative elements decreases going from left to right across a period.

The size of atoms of transition elements within the same period changes only slightly because electrons are filling \(d\) orbitals rather than the outermost energy level. Because the increase in nuclear charge is canceled by an increase in \(d\) electrons, the attraction of the valence electrons by the nucleus remains about the same. Because there is little change in the nuclear attraction for the valence electrons, the atomic size remains relatively constant for the transition elements.

**Sample Problem 5.10 Sizes of Atoms**

Identify the smaller atom in each of the following pairs:

a. N or F  
b. K or Kr  
c. Ca and Sr

**Solution**

a. The F atom has a greater positive charge on the nucleus, which pulls electrons closer, and makes the F atom smaller than the N atom. Atomic size decreases going from left to right across a period.

b. The Kr atom has a greater positive charge on the nucleus, which pulls electrons closer, and makes the Kr atom smaller than the K atom. Atomic size decreases going from left to right across a period.

c. The outer electrons in the Ca atom are closer to the nucleus than in the Sr atom, which makes the Ca atom smaller than the Sr atom. Atomic size increases going down a group.

**Study Check 5.10**

Which atom has the largest atomic size, P, As, or Se?

**Answer**

As
In an atom, negatively charged electrons are attracted to the positive charge of the protons in the nucleus. Therefore, energy is required to remove an electron from an atom. The ionization energy is the energy needed to remove one electron from an atom in the gaseous (g) state. When an electron is removed from a neutral atom, a cation with a 1+ charge is formed.

\[
\text{Na}(g) + \text{energy (ionization)} \rightarrow \text{Na}^+(g) + e^-
\]

The attraction of a nucleus for the outermost electrons decreases as those electrons are farther from the nucleus. Thus the ionization energy decreases going down a group (see FIGURE 5.12). However, going across a period from left to right, the positive charge of the nucleus increases because there is an increase in the number of protons. Thus the ionization energy increases going from left to right across the periodic table.

In summary, the ionization energy is low for the metals and high for the nonmetals. The high ionization energies of the noble gases indicate that their electron configurations are especially stable.

**SAMPLE PROBLEM 5.11 Ionization Energy**

Indicate the element in each set that has the higher ionization energy and explain your choice.

a. K or Na  
b. Mg or Cl  
c. F, N, or C

**TRY IT FIRST**

**SOLUTION**

a. Na. In Na, an electron is removed from a sublevel closer to the nucleus, which requires a higher ionization energy for Na compared with K.

b. Cl. The increased nuclear charge of Cl increases the attraction for the valence electrons, which requires a higher ionization energy for Cl compared to Mg.

c. F. The increased nuclear charge of F increases the attraction for the valence electrons, which requires a higher ionization energy for F compared to C or N.

**STUDY CHECK 5.11**

Arrange Sr, I, and Sn in order of increasing ionization energy.

**ANSWER**

Ionization energy increases going from left to right across a period: Sr, Sn, I.

**Metallic Character**

An element that has **metallic character** is an element that loses valence electrons easily. Metallic character is more prevalent in the elements on the left side of the periodic table (metals) and decreases going from left to right across a period. The elements on the right side of the periodic table (nonmetals) do not easily lose electrons, which means they are less metallic. Most of the metalloids between the metals and nonmetals tend to lose electrons, but not as easily as the metals. Thus, in Period 3, sodium, which loses electrons most easily, would be the most metallic. Going across from left to right in Period 3, metallic character decreases to argon, which has the least metallic character.

For elements in the same group of representative elements, metallic character increases going from top to bottom. Atoms at the bottom of any group have more electron levels, which makes it easier to lose electrons. Thus, the elements at the bottom of a group
on the periodic table have lower ionization energy and are more metallic compared to the elements at the top.

A summary of the trends in periodic properties we have discussed is given in TABLE 5.5.

<table>
<thead>
<tr>
<th>Periodic Property</th>
<th>Top to Bottom within a Group</th>
<th>Left to Right across a Period</th>
</tr>
</thead>
<tbody>
<tr>
<td>Valence Electrons</td>
<td>Remains the same</td>
<td>Increases</td>
</tr>
<tr>
<td>Atomic Size</td>
<td>Increases because there is an increase in the number of energy levels</td>
<td>Decreases as the number of protons increases, which strengthens the attraction of the nucleus for the valence electrons, and pulls them closer to the nucleus</td>
</tr>
<tr>
<td>Ionization Energy</td>
<td>Decreases because the valence electrons are easier to remove when they are farther from the nucleus</td>
<td>Increases as the number of protons increases, which strengthens the attraction of the nucleus for the valence electrons, and more energy is needed to remove a valence electron</td>
</tr>
<tr>
<td>Metallic Character</td>
<td>Increases because the valence electrons are easier to remove when they are farther from the nucleus</td>
<td>Decreases as the number of protons increases, which strengthens the attraction of the nucleus for the valence electrons, and makes it more difficult to remove a valence electron</td>
</tr>
</tbody>
</table>

**QUESTIONS AND PROBLEMS**

**5.6 Trends in Periodic Properties**

**LEARNING GOAL** Use the electron configurations of elements to explain the trends in periodic properties.

5.51 What do the group numbers from 1A (1) to 8A (18) for the elements indicate about electron configurations of those elements?

5.52 What is similar and what is different about the valence electrons of the elements in a group?

5.53 Write the group number using both A/B and 1 to 18 notations for elements that have the following outer electron configuration:

- a. \(2s^2\)
- b. \(3s^23p^3\)
- c. \(4s^23d^5\)
- d. \(5s^24d^{10}5p^4\)

5.54 Write the group number using both A/B and 1 to 18 notations for elements that have the following outer electron configuration:

- a. \(4s^23d^{10}4p^3\)
- b. \(4s^3\)
- c. \(4s^23d^8\)
- d. \(5s^24d^{10}5p^5\)

5.55 Write the valence electron configuration for each of the following:

- a. alkali metals
- b. Group 4A (14)
- c. Group 7A (17)
- d. Group 5A (15)

5.56 Write the valence electron configuration for each of the following:

- a. halogens
- b. Group 6A (16)
- c. Group 13
- d. alkaline earth metals

5.57 Indicate the number of valence electrons in each of the following:

- a. aluminum
- b. Group 5A (15)
- c. barium
- d. F, Cl, Br, and I

5.58 Indicate the number of valence electrons in each of the following:

- a. Li, Na, K, Rb, and Cs
- b. Se
- c. C, Si, Ge, Sn, and Pb
- d. Group 8A (18)

5.59 Write the group number and draw the Lewis symbol for each of the following elements:

- a. sulfur
- b. nitrogen
- c. calcium
- d. sodium
- e. gallium

5.60 Write the group number and draw the Lewis symbol for each of the following elements:

- a. carbon
- b. oxygen
- c. argon
- d. lithium
- e. chlorine

5.61 Select the larger atom in each pair.

- a. Na or Cl
- b. Na or Rb
- c. Na or Mg
- d. Rb or I

5.62 Select the larger atom in each pair.

- a. S or Ar
- b. S or O
- c. S or K
- d. S or Mg

5.63 Place the elements in each set in order of decreasing atomic size.

- a. Al, Si, Mg
- b. Cl, I, Br
- c. Sb, Sr, I
- d. P, Si, Na
5.64 Place the elements in each set in order of decreasing atomic size.
   a. Cl, S, P  
   b. Ge, Si, C  
   c. Ba, Ca, Sr  
   d. S, O, Se

5.65 Select the element in each pair with the higher ionization energy.
   a. Br or I  
   b. Mg or Sr  
   c. Si or P  
   d. I or Xe

5.66 Select the element in each pair with the higher ionization energy.
   a. Br or I  
   b. Mg or Sr  
   c. Si or P  
   d. I or Xe

5.67 Arrange each set of elements in order of increasing ionization energy.
   a. F, Cl, Br  
   b. Na, Cl, Al  
   c. Na, K, Cs  
   d. As, Ca, Br

5.68 Arrange each set of elements in order of increasing ionization energy.
   a. O, N, C  
   b. S, P, Cl  
   c. P, As, N  
   d. Al, Si, P

5.69 Place the following in order of decreasing metallic character:
   Br, Ge, Ca, Ga

5.70 Place the following in order of increasing metallic character:
   Na, P, Al, Ar

5.71 Fill in each of the following blanks using higher or lower, more or less: Sr has a _________ ionization energy and is _________ metallic than Sb.

5.72 Fill in each of the following blanks using higher or lower, more or less: N has a _________ ionization energy and is _________ metallic than As.

5.73 Complete each of the following statements a to d using 1, 2, or 3:
   1. decreases  
   2. increases  
   3. remains the same

5.74 Complete each of the following statements a to d using 1, 2, or 3:
   a. the ionization energy _________
   b. the atomic size _________
   c. the metallic character _________
   d. the number of valence electrons _________

5.75 Which statements completed with a to e will be true and which will be false?
   In Period 2, an atom of N compared to an atom of Li has a larger (greater)
   a. atomic size  
   b. ionization energy  
   c. number of protons  
   d. metallic character  
   e. number of valence electrons

5.76 Which statements completed with a to e will be true and which will be false?
   In Group 4A (14), an atom of C compared to an atom of Sn has a larger (greater)
   a. atomic size  
   b. ionization energy  
   c. number of protons  
   d. metallic character  
   e. number of valence electrons

5.77 a. What is the atomic number of In?
   b. How many electrons are in an atom of In?
   c. Use the sublevel blocks on the periodic table to write the electron configuration and abbreviated electron configuration for an atom of In.
   d. Write the group number and draw the Lewis symbol for In.
   e. Which is larger, an atom of indium or an atom of iodine?
   f. Which has a higher ionization energy, an atom of iodine or an atom of indium?

5.78 a. What is the atomic number of Te?
   b. How many electrons are in an atom of Te?
   c. Use the sublevel blocks on the periodic table to write the electron configuration and abbreviated electron configuration for an atom of Te.
   d. Write the group number and draw the Lewis symbol for Te.
   e. Which is smaller, an atom of selenium or an atom of tellurium?
   f. Which has a lower ionization energy, an atom of selenium or an atom of tellurium?
5.1 Electromagnetic Radiation

**LEARNING GOAL** Compare the wavelength, frequency, and energy of electromagnetic radiation.

- Electromagnetic radiation such as radio waves and visible light is energy that travels at the speed of light.
- Each particular type of radiation has a specific wavelength and frequency.
- A wavelength (symbol \( \lambda \), lambda) is the distance between a crest or trough in a wave and the next crest or trough on that wave.
- The frequency (symbol \( \nu \), nu) is the number of waves that pass a certain point in 1 s.
- All electromagnetic radiation travels at the speed of light (\( c \)), which is \( 3.00 \times 10^8 \) m/s.
- Mathematically, the relationship of the speed of light, wavelength, and frequency is expressed as \( c = \lambda \nu \).
- Long-wavelength radiation has low frequencies, while short-wavelength radiation has high frequencies.
- Radiation with a high frequency has high energy.
5.2 Atomic Spectra and Energy Levels

LEARNING GOAL Explain how atomic spectra correlate with the energy levels in atoms.
- The atomic spectra of elements are related to the specific energy levels occupied by electrons.
- Light consists of photons, which are particles of a specific energy.
- When an electron absorbs a photon of a particular energy, it attains a higher energy level. When an electron drops to a lower energy level, a photon of a particular energy is emitted.
- Each element has its own unique atomic spectrum.

5.3 Sublevels and Orbitals

LEARNING GOAL Describe the sublevels and orbitals for the electrons in an atom.
- An orbital is a region around the nucleus where an electron with a specific energy is most likely to be found.
- Each orbital holds a maximum of two electrons, which must have opposite spins.
- In each energy level (n), electrons occupy orbitals of identical energy within sublevels.
- An s sublevel contains one s orbital, a p sublevel contains three p orbitals, a d sublevel contains five d orbitals, and an f sublevel contains seven f orbitals.
- Each type of orbital has a unique shape.

5.4 Orbital Diagrams and Electron Configurations

LEARNING GOAL Draw the orbital diagram and write the electron configuration for an element.
- Within a sublevel, electrons enter orbitals in the same energy level one at a time until all the orbitals are half-filled.
- Additional electrons enter with opposite spins until the orbitals in that sublevel are filled with two electrons each.
- The orbital diagram for an element such as silicon shows the orbitals that are occupied by paired and unpaired electrons:

<table>
<thead>
<tr>
<th>1s</th>
<th>2s</th>
<th>2p</th>
<th>3s</th>
<th>3p</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td>↑</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
- The electron configuration for an element such as silicon shows the number of electrons in each sublevel: 1s²2s²2p⁶3s²3p².
- An abbreviated electron configuration for an element such as silicon places the symbol of a noble gas in brackets to represent the filled sublevels: [Ne]3s²3p².

5.5 Electron Configurations and the Periodic Table

LEARNING GOAL Write the electron configuration for an atom using the sublevel blocks on the periodic table.
- The periodic table consists of s, p, d, and f sublevel blocks.
- An electron configuration can be written following the order of the sublevel blocks on the periodic table.
- Beginning with 1s, an electron configuration is obtained by writing the sublevel blocks in order going across each period on the periodic table until the element is reached.

5.6 Trends in Periodic Properties

LEARNING GOAL Use the electron configurations of elements to explain the trends in periodic properties.
- The properties of elements are related to the valence electrons of the atoms.
- With only a few exceptions, each group of elements has the same arrangement of valence electrons differing only in the energy level.
- Valence electrons are represented as dots around the symbol of the element in the Lewis symbol.
- The size of an atom increases going down a group and decreases going from left to right across a period.
- The energy required to remove a valence electron is the ionization energy, which decreases going down a group, and increases going from left to right across a period.
- The metallic character of an element increases going down a group and decreases going from left to right across a period.

KEY TERMS

atomic size The distance between the outermost electrons and the nucleus.
atomic spectrum A series of lines specific for each element produced by photons emitted by electrons dropping to lower energy levels.
d block The 10 elements in Groups 3B (3) to 2B (12) in which electrons fill the five d orbitals.
electromagnetic radiation Forms of energy such as visible light, microwaves, radio waves, infrared, ultraviolet light, and X-rays that travel as waves at the speed of light.
electromagnetic spectrum The arrangement of types of radiation from long wavelengths to short wavelengths.
electron configuration A list of the number of electrons in each sublevel within an atom, arranged by increasing energy.
energy level A group of electrons with similar energy.

f block The 14 elements in the rows at the bottom of the periodic table in which electrons fill the seven 4f and 5f orbitals.

frequency The number of times the crests of a wave pass a point in 1 s.

ionization energy The energy needed to remove the least tightly bound electron from the outermost energy level of an atom.

Lewis symbol The representation of an atom that shows valence electrons as dots around the symbol of the element.

metallic character A measure of how easily an element loses a valence electron.

orbital The region around the nucleus of an atom where electrons of certain energy are most likely to be found: s orbitals are spherical; p orbitals have two lobes.

orbital diagram A diagram that shows the distribution of electrons in the orbitals of the energy levels.

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

Writing Electron Configurations (5.4)
• The electron configuration for an atom specifies the energy levels and sublevels occupied by the electrons of an atom.
• An electron configuration is written starting with the lowest energy sublevel, followed by the next lowest energy sublevel.
• The number of electrons in each sublevel is shown as a superscript.

Example: Write the electron configuration for palladium.
Answer: Palladium has atomic number 46, which means it has 46 protons and 46 electrons.
\[
1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}
\]

Using the Periodic Table to Write Electron Configurations (5.5)
• An electron configuration corresponds to the location of an element on the periodic table, where different blocks within the periodic table are identified as the s, p, d, and f sublevels.

Example: Use the periodic table to write the electron configuration for sulfur.
Answer: Sulfur (atomic number 16) is in Group 6A (16) and Period 3.

<table>
<thead>
<tr>
<th>Period</th>
<th>Sublevel Block Filling</th>
<th>Sublevel Block Notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1s (H → He)</td>
<td>1s²</td>
</tr>
<tr>
<td>2</td>
<td>2s (Li → Be) and 2p (B → Ne)</td>
<td>2s² 2p⁶</td>
</tr>
<tr>
<td>3</td>
<td>3s (Na → Mg) and 3p (Al → S)</td>
<td>3s² 3p⁵</td>
</tr>
</tbody>
</table>

The electron configuration for sulfur (S) is: 1s² 2s² 2p⁶ 3s² 3p⁴.

Identifying Trends in Periodic Properties (5.6)
• The size of an atom increases going down a group and decreases going from left to right across a period.
• The ionization energy decreases going down a group and increases going from left to right across a period.
• The metallic character of an element increases going down a group and decreases going from left to right across a period.

Example: For Mg, P, and Cl, identify which has the
a. largest atomic size
b. highest ionization energy
c. greatest metallic character

Answer: a. Mg    b. Cl    c. Mg

Drawing Lewis Symbols (5.6)
• The valence electrons are the electrons in the s and p sublevels with the highest principal quantum number n.
• The number of valence electrons is the same as the group number for the representative elements.
• A Lewis symbol represents the number of valence electrons as dots placed around the symbol for the element.

Example: Give the group number and number of valence electrons, and draw the Lewis symbol for each of the following:

a. Rb
b. Se
c. Xe

Answer: a. Group 1A (1), one valence electron, Rb
b. Group 6A (16), six valence electrons, Se
  c. Group 8A (18), eight valence electrons, Xe
**UNDErstanding the Concepts**

*The chapter sections to review are shown in parentheses at the end of each question.*

Use the following diagram for problems 5.79 and 5.80:

A.

B.

C.

5.79 Select diagram A, B, or C that (5.1)
- a. has the longest wavelength
- b. has the shortest wavelength
- c. has the highest frequency
- d. has the lowest frequency

5.80 Select diagram A, B, or C that (5.1)
- a. has the highest energy
- b. has the lowest energy
- c. would represent blue light
- d. would represent red light

5.81 Match the following with an s or p orbital: (5.3)
- a. 
- b. 
- c. 

5.82 Match the following with s or p orbitals: (5.3)
- a. two lobes
- b. spherical shape
- c. found in \( n = 1 \)
- d. found in \( n = 3 \)

**ADDITIONAL QUESTIONS AND PROBLEMS**

5.87 What is the difference between a continuous spectrum and an atomic spectrum? (5.1)

5.88 Why does a neon sign give off red light? (5.1)

5.89 What is the Pauli exclusion principle? (5.3)

5.90 Why would there be five unpaired electrons in a d sublevel but no paired electrons? (5.3)

5.91 Which of the following orbitals are not possible in an atom: 1d, 2s, 3p, and 5f? (5.3)

5.92 Which of the following orbitals are not possible in an atom: 3f, 4p, 5d, and 6p? (5.3)

5.83 Indicate whether or not the following orbital diagrams are possible and explain. When possible, indicate the element it represents. (5.4)
- a. 
- b. 

5.84 Indicate whether or not the following abbreviated orbital diagrams are possible and explain. When possible, indicate the element it represents. (5.4)
- a. 
- b. 

5.85 Match the spheres A through D with atoms of B, Al, Ga, and In. (5.6)

5.86 Match the spheres A through D with atoms of Na, Mg, P, and Cl. (5.6)

5.93 a. What electron sublevel starts to fill after completion of the 3s sublevel? (5.4)
- b. What electron sublevel starts to fill after completion of the 4p sublevel?
- c. What electron sublevel starts to fill after completion of the 3d sublevel?
- d. What electron sublevel starts to fill after completion of the 3p sublevel?

5.94 a. What electron sublevel starts to fill after completion of the 5s sublevel? (5.4)
- b. What electron sublevel starts to fill after completion of the 4d sublevel?
- c. What electron sublevel starts to fill after completion of the 4f sublevel?
- d. What electron sublevel starts to fill after completion of the 5p sublevel?
5.95  a. How many 3d electrons are in Fe? (5.4)  
b. How many 5p electrons are in Ba?  
c. How many 4d electrons are in I?  
d. How many 7s electrons are in Ra?

5.96  a. How many 4d electrons are in Cd? (5.4)  
b. How many 4p electrons are in Br?  
c. How many 6p electrons are in Bi?  
d. How many 4s electrons are in Zn?

5.97  What do the elements Ca, Sr, and Ba have in common in terms of their electron configuration? Where are they located on the periodic table? (5.4, 5.5)

5.98  What do the elements O, S, and Se have in common in terms of their electron configuration? Where are they located on the periodic table? (5.4, 5.5)

5.99  Consider three elements with the following abbreviated electron configurations: (5.4, 5.5, 5.6)

\[ \text{X} = [\text{Ar}]4s^2 \]  
\[ \text{Y} = [\text{Ne}]3s^23p^6 \]  
\[ \text{Z} = [\text{Ar}]4s^23d^{10}4p^5 \]

a. Identify each element as a metal, nonmetal, or metalloid.  
b. Which element has the largest atomic size?  
c. Which element has the highest ionization energy?  
d. Which element has the smallest atomic size?

5.100  Consider three elements with the following abbreviated electron configurations: (5.4, 5.5, 5.6)

\[ \text{X} = [\text{Ar}]4s^23d^5 \]  
\[ \text{Y} = [\text{Ar}]4s^23d^{10}4p^1 \]  
\[ \text{Z} = [\text{Ar}]4s^23d^{10}4p^6 \]

a. Identify each element as a metal, nonmetal, or metalloid.  
b. Which element has the smallest atomic size?  
c. Which element has the highest ionization energy?  
d. Which element has a half-filled sublevel?

5.101  Name the element that corresponds to each of the following: (5.4, 5.5, 5.6)

a. \(1s^22s^22p^63s^23p^3\)  
b. alkali metal with the smallest atomic size  
c. \([\text{Kr}]5s^24d^{10}\)  
d. Group 5A (15) element with the highest ionization energy  
e. Period 3 element with the largest atomic size

5.102  Name the element that corresponds to each of the following: (5.4, 5.5, 5.6)

a. \(1s^22s^22p^63s^23p^64s^23d^5\)  
b. \([\text{Xe}]6s^24f^{14}5d^{10}6p^3\)  
c. halogen with the highest ionization energy  
d. Group 2A (2) element with the lowest ionization energy  
e. Period 4 element with the smallest atomic size

5.103  Why is the ionization energy of Ca higher than that of K but lower than that of Mg? (5.6)

5.104  Why is the ionization energy of Br lower than that of Cl but higher than that of Se? (5.6)

5.105  Select the element with higher ionization energy in each pair. (5.6)

a. Cs or Ba  
b. Cl or At  
c. Si or P  
d. Be or Ca

5.106  Select the more metallic element in each pair. (5.6)

a. P or Bi  
b. S or Te  
c. Rb or In  
d. Li or Ne

5.107  Of the elements Mg, Cu, Si, and Cl, which (5.6)

a. is a semiconductor?  
b. is in Group 2A (2)?  
c. has the highest ionization energy?  
d. has the smallest atomic size?  
e. is found in Group 7A (17)?

5.109  Write the abbreviated electron configuration and group number for each of the following elements: (5.4)

a. Zn  
b. I  
c. V  
d. Sr

5.110  Write the abbreviated electron configuration and group number for each of the following elements: (5.4)

a. Cu  
b. Fe  
c. P  
d. As

5.111  Write the group number and draw the Lewis symbol for each of the following elements: (5.6)

a. barium  
b. fluorine  
c. krypton  
d. arsenic

5.112  Write the group number and draw the Lewis symbol for each of the following elements: (5.6)

a. neon  
b. iodine  
c. bismuth  
d. tin

5.115  How do scientists explain the colored lines observed in the spectra of heated atoms? (5.2)

5.116  Even though H has only one electron, there are many lines in the atomic spectrum of H. Explain. (5.2)

5.117  What is meant by an energy level, a sublevel, and an orbital? (5.3)

5.118  In some periodic tables, H is placed in Group 1A (1). In other periodic tables, H is also placed in Group 7A (17). Why? (5.4, 5.5)

5.119  Compare Ar, S, and Se in terms of atomic size and ionization energy. (5.6)

5.120  Compare B, O, and Al in terms of atomic size and ionization energy. (5.6)

**CHALLENGE QUESTIONS**

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

5.113  Give the symbol of the element that has the (5.6)

a. smallest atomic size in Group 6A (16)  
b. smallest atomic size in Period 3  
c. highest ionization energy in Group 3A (13)  
d. lowest ionization energy in Period 3  
e. abbreviated electron configuration \([\text{Kr}]5s^24d^9\)

5.114  Give the symbol of the element that has the (5.6)

a. largest atomic size in Period 5  
b. largest atomic size in Group 2A (2)  
c. highest ionization energy in Group 8A (18)  
d. lowest ionization energy in Period 2  
e. abbreviated electron configuration \([\text{Kr}]5s^24d^{10}5p^2\)
5.21 a. spherical 
   b. two lobes 
   c. spherical

5.23 a. 1 and 2 
   b. 3 
   c. 1 and 2 
   d. 1, 2, and 3

5.25 a. There are five orbitals in the 3d sublevel. 
   b. There is one sublevel in the n = 1 energy level. 
   c. There is one orbital in the 6s sublevel. 
   d. There are nine orbitals in the n = 3 energy level.

5.27 a. There is a maximum of two electrons in a 2p orbital. 
   b. There is a maximum of six electrons in the 3p sublevel. 
   c. There is a maximum of 32 electrons in the n = 4 energy level. 
   d. There is a maximum of 10 electrons in the 5d sublevel.

5.29 The electron configuration shows the number of electrons in each sublevel of an atom. The abbreviated electron configuration uses the symbol of the preceding noble gas to show completed sublevels.

5.31 a. 1s 2s 2p 
   b. 1s 2s 2p 3s 3p 
   c. 1s 2s 2p 3s 3p 4s 3d 
   d. 1s 2s 2p 3s 3p 4s 3d 5p

5.33 a. 1s2 2s2 2p6 3s2 3p6 4s2 3d8 
   b. 1s2 2s2 2p6 3s2 
   c. 1s2 2s2 
   d. 1s2 2s2 2p6 3s2 3p6 4s2 3d2

5.35 a. [Kr]5s2 4d10 5p2 
   b. [Kr]5s2 4d10 
   c. [Ar]4s2 3d10 4p6 
   d. [He]2s2 2p5

5.37 a. Li 
   b. Ti 
   c. Ge 
   d. F

5.39 a. Al 
   b. C 
   c. Ar 
   d. Be

5.41 a. 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p3 
   b. 1s2 2s2 2p6 3s2 3p6 4s2 3d6 
   c. 1s2 2s2 2p6 3s2 3p6 4s2 3d10 
   d. 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6

5.43 a. [Ar]4s2 3d2 
   b. [Ar]4s2 3d10 4p5 
   c. [Xe]6s2 
   d. [Xe]6s2 4f14 5d106p2

5.45 a. P 
   b. Co 
   c. Zn 
   d. Bi

5.47 a. Ga 
   b. N 
   c. Xe 
   d. Zr

5.49 a. 10 
   b. 6 
   c. 3 
   d. 1

5.51 The group numbers 1A to 8A indicate the number of valence electrons from 1 to 8.

5.53 a. 2A (2) 
   b. 5A (15) 
   c. 7B (7) 
   d. 6A (16)

5.55 a. ns1 
   b. ns2 np3 
   c. ns2 np5 
   d. ns2 np6

5.57 a. 3 
   b. 5 
   c. 2 
   d. 7

5.59 a. Sulfur is in Group 6A (16); S
   b. Nitrogen is in Group 5A (15); N
   c. Calcium is in Group 2A (2); Ca
   d. Sodium is in Group 1A (1); Na
   e. Gallium is in Group 3A (13); Ga

5.61 a. Na 
   b. Rb 
   c. Na 
   d. Rb

5.63 a. Mg, Al, Si 
   b. I, Br, Cl 
   c. Sr, Sb, I 
   d. Na, Si, P

5.65 a. Br 
   b. Mg 
   c. P 
   d. Xe

5.67 a. Br, Cl, F 
   b. Na, Al, Cl 
   c. Cs, K, Na 
   d. Ca, As, Br

5.69 Ca, Ga, Ge, Br

5.71 lower, more

5.73 a. 1. decreases 
   b. 2. increases 
   c. 2. increases 
   d. 3. remains the same

5.75 a. false 
   b. true 
   c. true 
   d. false 
   e. true

5.77 a. 49 
   b. 49 
   c. 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d10 5p1; [Kr]5s2 4d10 5p1 
   d. 3A (13) 
   e. indium 
   f. iodine

5.79 a. C has the longest wavelength. 
   b. A has the shortest wavelength. 
   c. A has the highest frequency. 
   d. C has the lowest frequency.

5.81 a. p 
   b. s 
   c. p

5.83 a. This is possible. This element is magnesium. 
   b. Not possible. The 2p sublevel would fill before the 3s, and only two electrons are allowed in an s orbital.

5.85 B is D, Al is A, Ga is C, and In is B.

5.87 A continuous spectrum from white light contains wavelengths of all energies. Atomic spectra are line spectra in which a series of lines corresponds to energy emitted when electrons drop from a higher energy level to a lower level.

5.89 The Pauli exclusion principle states that two electrons in the same orbital must have opposite spins.
5.91 A 1d orbital is not possible because \( n = 1 \) energy level has only an s sublevel. A 2s orbital is possible because the \( n = 2 \) energy level has two sublevels, including s and p sublevels. A 3p orbital is possible because the \( n = 3 \) energy level has three sublevels, including s, p, and d sublevels. A 5f sublevel is possible in the \( n = 5 \) energy level because five sublevels are allowed, including a 5f sublevel.

5.93 a. 3p  b. 5s  c. 4p  d. 4s

5.97 Ca, Sr, and Ba all have two valence electrons, \( ns^2 \), which place them in Group 2A (2).

5.99 a. X is a metal; Y and Z are nonmetals.
   b. X has the largest atomic size.
   c. Y has the highest ionization energy.
   d. Y has the smallest atomic size.

5.101 a. phosphorus  b. lithium (H is a nonmetal)
   c. cadmium  d. nitrogen
   e. sodium

5.103 Calcium has a greater number of protons than K. The least tightly bound electron in Ca is farther from the nucleus than in Mg and needs less energy to remove.


5.107 a. Si  b. Mg  c. Cl
d. Cl  e. Cu

5.109 a. [Ar]4s^23d^10; Group 2B (12)
   b. [Kr]5s^24d^105p^4; Group 7A (17)
   c. [Ar]4s^23d^3; Group 5B (5)
   d. [Kr]5s^2; Group 2A (2)

5.111 a. Barium in Group 2A (2); Ba
   b. Fluorine is in Group 7A (17); F
   c. Krypton is in Group 8A (18); Kr
   d. Arsenic is in Group 5A (15); As

5.113 a. O  b. Ar  c. B
d. Na  e. Ru

5.115 The series of lines separated by dark sections in atomic spectra indicate that the energy emitted by the elements is not continuous and that electrons are moving between discrete energy levels.

5.117 The energy level contains all the electrons with similar energy. A sublevel contains electrons with the same energy, while an orbital is the region around the nucleus where electrons of a certain energy are most likely to be found.

5.119 Se has a larger atomic size than S; Ar is smaller than S:
Ar < S < Se. Ar has the highest ionization energy; S has a higher ionization energy than Se: Ar > S > Se.
RICHARD’S DOCTOR HAS recommended that he take a low-dose aspirin (81 mg) every day to prevent a heart attack or stroke. Richard is concerned about taking aspirin and asks Sarah, a pharmacist working at a local pharmacy, about the effects of aspirin. Sarah explains to Richard that aspirin is acetylsalicylic acid and has the chemical formula $C_9H_8O_4$. Aspirin is a molecular compound, often referred to as an organic molecule because it contains the nonmetals carbon (C), hydrogen (H), and oxygen (O). Sarah explains to Richard that aspirin is used to relieve minor pains, to reduce inflammation and fever, and to slow blood clotting. Aspirin is one of several nonsteroidal anti-inflammatory drugs (NSAIDs) that reduce pain and fever by blocking the formation of prostaglandins, which are chemical messengers that transmit pain signals to the brain and cause fever. Some potential side effects of aspirin may include heartburn, upset stomach, nausea, and an increased risk of a stomach ulcer.

CAREER
Pharmacist
Pharmacists work in hospitals, pharmacies, clinics, and long-term care facilities where they are responsible for the preparation and distribution of pharmaceutical medications based on a doctor’s orders. They obtain the proper medication, and also calculate, measure, and label the patient’s medication. Pharmacists advise clients and health care practitioners on the selection of both prescription and over-the-counter drugs, proper dosages, and geriatric considerations, as well as possible side effects and interactions. They may also administer vaccinations; prepare sterile intravenous solutions; and advise clients about health, diet, and home medical equipment. Pharmacists also prepare insurance claims, and create and maintain patient profiles.
Most of the elements, except the noble gases, are found in nature combined as compounds. The noble gases are so stable that they form compounds only under extreme conditions. One explanation for the stability of noble gases is that they have a filled valence electron energy level.

Compounds form when electrons are transferred or shared to give stable electron configuration to the atoms. In the formation of either an ionic bond or a covalent bond, atoms lose, gain, or share valence electrons to acquire an octet of eight valence electrons. This tendency of atoms to attain a stable electron configuration is known as the octet rule and provides a key to our understanding of the ways in which atoms bond and form compounds. A few elements achieve the stability of helium with two valence electrons. However, we do not use the octet rule with transition elements.

Ionic bonds occur when the valence electrons of atoms of a metal are transferred to atoms of nonmetals. For example, sodium atoms lose electrons and chlorine atoms gain electrons to form the ionic compound NaCl. Covalent bonds form when atoms of nonmetals share valence electrons. In the molecular compounds H₂O and C₃H₈, atoms share electrons.

### 6.1 Ions: Transfer of Electrons

**LEARNING GOAL** Write the symbols for the simple ions of the representative elements.

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<table>
<thead>
<tr>
<th>Type</th>
<th>Ionic Compounds</th>
<th>Molecular Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Particles</td>
<td>Ions</td>
<td>Molecules</td>
</tr>
<tr>
<td>Bonding</td>
<td>Ionic bonds</td>
<td>Covalent bonds</td>
</tr>
<tr>
<td>Examples</td>
<td>Na⁺ Cl⁻ ions</td>
<td>H₂O molecules</td>
</tr>
<tr>
<td></td>
<td>C₃H₈ molecules</td>
<td></td>
</tr>
</tbody>
</table>

**Positive Ions: Loss of Electrons**

In ionic bonding, **ions**, which have electrical charges, form when atoms lose or gain electrons to form a stable electron configuration. Because the ionization energies of metals of Groups 1A (1), 2A (2), and 3A (13) are low, metal atoms readily lose their valence electrons. In doing so, they form ions with positive charges. A metal atom obtains the same electron configuration as its nearest noble gas (usually eight valence electrons). For example, when a sodium atom loses its single valence electron, the remaining electrons have a stable electron configuration. By losing an electron, sodium has 10 negatively charged electrons instead of 11. Because there are still 11 positively charged protons in its nucleus, the atom is no longer neutral. It is now a sodium ion with a positive electrical charge, called an **ionic charge**, of 1+. In the Lewis symbol for the sodium ion, the ionic charge of...
1+ is written in the upper right-hand corner, Na+, where the 1 is understood. The sodium ion is smaller than the sodium atom because the ion has lost its outermost electron from the third energy level. A positively charged ion of a metal is called a cation (pronounced cat-eye-un) and uses the name of the element.

\[
\text{Ionic charge} = \text{Charge of protons} + \text{Charge of electrons} \\
1+ = (11+) + (10-) 
\]

### Ionic Charge Examples

<table>
<thead>
<tr>
<th>Name</th>
<th>Lewis Symbol</th>
<th>Protons</th>
<th>Electrons</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium atom</td>
<td>Na</td>
<td>11</td>
<td>11</td>
<td>1s²2s²2p⁶3s¹</td>
</tr>
<tr>
<td>Sodium ion</td>
<td>Na⁺</td>
<td>11</td>
<td>10</td>
<td>1s²2s²2p⁶</td>
</tr>
</tbody>
</table>

Magnesium, a metal in Group 2A (2), obtains a stable electron configuration by losing two valence electrons to form a magnesium ion with a 2⁺ ionic charge, Mg²⁺. The magnesium ion is smaller than the magnesium atom because the outermost electrons in the third energy level were removed. The octet in the magnesium ion is made up of electrons that fill its second energy level.

### Electron Configuration Examples

<table>
<thead>
<tr>
<th>Name</th>
<th>Lewis Symbol</th>
<th>Protons</th>
<th>Electrons</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Magnesium atom</td>
<td>Mg</td>
<td>12</td>
<td>12</td>
<td>1s²2s²2p⁶3s²</td>
</tr>
<tr>
<td>Magnesium ion</td>
<td>Mg⁺⁺</td>
<td>12</td>
<td>10</td>
<td>1s²2s²2p⁶</td>
</tr>
</tbody>
</table>

#### Negative Ions: Gain of Electrons

The ionization energy of a nonmetal atom in Groups 5A (15), 6A (16), or 7A (17) is high. In an ionic compound, a nonmetal atom gains one or more valence electrons to obtain a stable electron configuration. By gaining electrons, a nonmetal atom forms a negatively charged ion. For example, an atom of chlorine with seven valence electrons gains one electron to form an octet. Because it now has 18 electrons and 17 protons in its nucleus, the chlorine atom is no longer neutral. It is a chloride ion with an ionic charge of 1⁻, which is written as Cl⁻, with the 1 understood. A negatively charged ion, called an anion (pronounced an-eye-un), is named by using the first syllable of its element name followed by ide. The chloride ion is larger than the chlorine atom because the ion has an additional electron, which completes its outermost energy level.

\[
\text{Ionic charge} = \text{Charge of protons} + \text{Charge of electrons} \\
1⁻ = (17+) + (18-) 
\]
### Table 6.1 Formulas and Names of Some Common Ions

<table>
<thead>
<tr>
<th>Group Number</th>
<th>Metals</th>
<th>Cation</th>
<th>Name of Cation</th>
<th>Group Number</th>
<th>Nonmetals</th>
<th>Anion</th>
<th>Name of Anion</th>
</tr>
</thead>
<tbody>
<tr>
<td>1A (1)</td>
<td>Li⁺</td>
<td>Lithium</td>
<td></td>
<td>5A (15)</td>
<td>N³⁻</td>
<td>Nitride</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Na⁺</td>
<td>Sodium</td>
<td></td>
<td></td>
<td>P³⁻</td>
<td>Phosphide</td>
<td></td>
</tr>
<tr>
<td>2A (2)</td>
<td>K⁺</td>
<td>Potassium</td>
<td></td>
<td>6A (16)</td>
<td>O²⁻</td>
<td>Oxide</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Mg²⁺</td>
<td>Magnesium</td>
<td></td>
<td></td>
<td>S²⁻</td>
<td>Sulfide</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Ca²⁺</td>
<td>Calcium</td>
<td></td>
<td>7A (17)</td>
<td>F⁻</td>
<td>Fluoride</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Br²⁺</td>
<td>Barium</td>
<td></td>
<td></td>
<td>Cl⁻</td>
<td>Chloride</td>
<td></td>
</tr>
<tr>
<td>3A (13)</td>
<td>Al³⁺</td>
<td>Aluminum</td>
<td></td>
<td></td>
<td>Br⁻</td>
<td>Bromide</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>I⁻</td>
<td>Iodide</td>
<td></td>
</tr>
</tbody>
</table>

**TABLE 6.1** lists the names of some important metal and nonmetal ions.

**SAMPLE PROBLEM 6.1 Ions**

a. Write the symbol and name for the ion that has 7 protons and 10 electrons.
b. Write the symbol and name for the ion that has 20 protons and 18 electrons.

**TRY IT FIRST**

**SOLUTION**

a. The element with 7 protons is nitrogen. In an ion of nitrogen with 10 electrons, the ionic charge would be 3⁻, \( [(7^+) + (10^-)] = 3^- \). The ion, written as N³⁻, is the nitride ion.
b. The element with 20 protons is calcium. In an ion of calcium with 18 electrons, the ionic charge would be 2⁺, \( [(20^+) + (18^-)] = 2^+ \). The ion, written as Ca²⁺, is the calcium ion.

**STUDY CHECK 6.1**

How many protons and electrons are in each of the following ions?

**a.** Sr²⁺  
**b.** Cl⁻

**ANSWER**

**a.** 38 protons, 36 electrons  
**b.** 17 protons, 18 electrons
Ionic Charges from Group Numbers

In ionic compounds, representative elements usually lose or gain electrons to give eight valence electrons like their nearest noble gas (or two for helium). We can use the group numbers in the periodic table to determine the charges for the ions of the representative elements. The elements in Group 1A (1) lose one electron to form ions with a 1+ charge. The elements in Group 2A (2) lose two electrons to form ions with a 2+ charge. The elements in Group 3A (13) lose three electrons to form ions with a 3+ charge. In this text, we do not use the group numbers of the transition elements to determine their ionic charges.

In ionic compounds, the elements in Group 7A (17) gain one electron to form ions with a 1− charge. The elements in Group 6A (16) gain two electrons to form ions with a 2− charge. The elements in Group 5A (15) gain three electrons to form ions with a 3− charge. The nonmetals of Group 4A (14) do not typically form ions. However, the metals Sn and Pb in Group 4A (14) lose electrons to form positive ions. TABLE 6.2 lists the ionic charges for some common monatomic ions of representative elements.

<table>
<thead>
<tr>
<th>Noble Gases</th>
<th>Metals Lose Valence Electrons</th>
<th>Nonmetals Gain Valence Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>1A (1)</td>
<td>2A (2)</td>
</tr>
<tr>
<td>He</td>
<td>Li+</td>
<td>Ne</td>
</tr>
<tr>
<td>Ne</td>
<td>Na+ Mg2+ Al3+</td>
<td>Ar</td>
</tr>
<tr>
<td>Ar</td>
<td>K+ Ca2+</td>
<td>Kr</td>
</tr>
<tr>
<td>Kr</td>
<td>Rb+ Sr2+</td>
<td>Xe</td>
</tr>
<tr>
<td>Xe</td>
<td>Cs+ Ba2+</td>
<td>At−</td>
</tr>
</tbody>
</table>

**TABLE 6.2** Examples of Monatomic Ions and Their Nearest Noble Gases

**SAMPLE PROBLEM 6.2** Writing Symbols for Ions

Consider the elements aluminum and oxygen.

a. Identify each as a metal or a nonmetal.
b. State the number of valence electrons for each.
c. State the number of electrons that must be lost or gained for each to achieve an octet.
d. Write the symbol, including its ionic charge, and the name for each resulting ion.

**SOLUTION**

<table>
<thead>
<tr>
<th>Aluminum</th>
<th>Oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. metal</td>
<td>nonmetal</td>
</tr>
<tr>
<td>b. three valence electrons</td>
<td>six valence electrons</td>
</tr>
<tr>
<td>c. loses 3 e−</td>
<td>gains 2 e−</td>
</tr>
<tr>
<td>d. Al3+, [(13+) + (10−) = 3+], aluminum ion</td>
<td>O2−, [(8+) + (10−) = 2−], oxide ion</td>
</tr>
</tbody>
</table>
STUDY CHECK 6.2
Write the symbols for the ions formed by potassium and sulfur.

ANSWER
K\(^+\) and S\(^{2-}\)

CHEMISTRY LINK TO HEALTH

Some Important Ions in the Body

Several ions in body fluids have important physiological and metabolic functions. Some are listed in TABLE 6.3.

Foods such as bananas, milk, cheese, and potatoes provide the body with ions that are important in regulating body functions.

Milk, cheese, bananas, cereal, and potatoes provide ions for the body.

<table>
<thead>
<tr>
<th>Ion</th>
<th>Occurrence</th>
<th>Function</th>
<th>Source</th>
<th>Result of Too Little</th>
<th>Result of Too Much</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na(^+)</td>
<td>Principal cation outside the cell</td>
<td>Regulation and control of body fluids</td>
<td>Salt, cheese, pickles</td>
<td>Hyponatremia, anxiety, diarrhea, circulatory failure, decrease in body fluid</td>
<td>Hypernatremia, little urine, thirst, edema</td>
</tr>
<tr>
<td>K(^+)</td>
<td>Principal cation inside the cell</td>
<td>Regulation of body fluids and cellular functions</td>
<td>Bananas, orange juice, milk, prunes, potatoes</td>
<td>Hypokalemia (hypopotassemia), lethargy, muscle weakness, failure of neurological impulses</td>
<td>Hyperkalemia (hyperpotassemia), irritability, nausea, little urine, cardiac arrest</td>
</tr>
<tr>
<td>Ca(^{2+})</td>
<td>Cation outside the cell; 90% of calcium in the body in bones</td>
<td>Major cation of bones; needed for muscle contraction</td>
<td>Milk, yogurt, cheese, greens, spinach</td>
<td>Hypocalcemia, tingling fingertips, muscle cramps, osteoporosis</td>
<td>Hypercalcemia, relaxed muscles, kidney stones, deep bone pain</td>
</tr>
<tr>
<td>Mg(^{2+})</td>
<td>Cation outside the cell; 50% of magnesium in the body in bone structure</td>
<td>Essential for certain enzymes, muscles, nerve control</td>
<td>Widely distributed (part of chlorophyll of all green plants), nuts, whole grains</td>
<td>Disorientation, hypertension, tremors, slow pulse</td>
<td>Drowsiness</td>
</tr>
<tr>
<td>Cl(^-)</td>
<td>Principal anion outside the cell</td>
<td>Gastric juice, regulation of body fluids</td>
<td>Salt</td>
<td>Same as for Na(^+)</td>
<td>Same as for Na(^+)</td>
</tr>
</tbody>
</table>
6.1 Ions: Transfer of Electrons

LEARNING GOAL Write the symbols for the simple ions of the representative elements.

6.1 State the number of electrons that must be lost by atoms of each of the following to achieve a stable electron configuration:
   a. Li  b. Ca  c. Ga  d. Cs  e. Ba
6.2 State the number of electrons that must be gained by atoms of each of the following to achieve a stable electron configuration:
   a. Cl  b. Se  c. N  d. I  e. S
6.3 State the number of electrons lost or gained when the following elements form ions:
6.4 State the number of electrons lost or gained when the following elements form ions:
   a. O  b. Group 2A (2)  c. F  d. K  e. Rb
6.5 Write the symbols for the ions with the following number of protons and electrons:
   a. 3 protons, 2 electrons  b. 9 protons, 10 electrons  c. 12 protons, 10 electrons  d. 26 protons, 23 electrons
6.6 Write the symbols for the ions with the following number of protons and electrons:
   a. 8 protons, 10 electrons  b. 19 protons, 18 electrons  c. 35 protons, 36 electrons  d. 50 protons, 46 electrons
6.7 State the number of protons and electrons in each of the following:
   a. Cu²⁺  b. Se²⁻  c. Br⁻  d. Fe³⁺
6.8 State the number of protons and electrons in each of the following:
   a. P³⁻  b. Ni²⁺  c. Au³⁺  d. Ag⁺
6.9 Write the symbol for the ion of each of the following:
   a. chlorine  b. cesium  c. nitrogen  d. radium
6.10 Write the symbol for the ion of each of the following:
   a. fluorine  b. calcium  c. sodium  d. iodine
6.11 Write the names for each of the following ions:
   a. Li⁺  b. Ca²⁺  c. Ga³⁺  d. P³⁻
6.12 Write the names for each of the following ions:
   a. Rb⁺  b. Sr²⁺  c. S²⁻  d. F⁻

Applications

6.13 State the number of protons and electrons in each of the following ions:
   a. O²⁻, used to build biomolecules and water  b. K⁺, most prevalent positive ion in cells, needed for muscle contraction, nerve impulses  c. I⁻, needed for thyroid function  d. Na⁺, most prevalent positive ion in extracellular fluid
6.14 State the number of protons and electrons in each of the following ions:
   a. P³⁻, needed for bones and teeth  b. F⁻, used to strengthen tooth enamel  c. Mg²⁺, needed for bones and teeth  d. Ca²⁺, needed for bones and teeth

6.2 Ionic Compounds

LEARNING GOAL Using charge balance, write the correct formula for an ionic compound.

We utilize ionic compounds such as salt, NaCl, and baking soda, NaHCO₃, every day. Milk of magnesia, Mg(OH)₂, or calcium carbonate, CaCO₃, may be taken to settle an upset stomach. In a mineral supplement, iron may be present as iron(II) sulfate, FeSO₄, iodine as potassium iodide, KI, and manganese as manganese(II) sulfate, MnSO₄. Some sunscreens contain zinc oxide, ZnO, while tin(II) fluoride, SnF₂, in toothpaste provides fluoride to help prevent tooth decay. Gemstones are ionic compounds that are cut and polished to make jewelry. For example, sapphires and rubies are aluminum oxide, Al₂O₃. Impurities of chromium ions make rubies red, and iron and titanium ions make sapphires blue.
Properties of Ionic Compounds

In an ionic compound, one or more electrons are transferred from metals to nonmetals, which form positive and negative ions. The attraction between these ions is called an ionic bond.

The physical and chemical properties of an ionic compound such as NaCl are very different from those of the original elements. For example, the original elements of NaCl were sodium, which is a soft, shiny metal, and chlorine, which is a yellow-green poisonous gas. However, when they react and form positive and negative ions, they produce NaCl, which is ordinary table salt, a hard, white, crystalline substance that is important in our diet.

In a crystal of NaCl, the larger Cl\(^{-}\) ions are arranged in a three-dimensional structure in which the smaller Na\(^{+}\) ions occupy the spaces between the Cl\(^{-}\) ions (see Figure 6.1). In this crystal, every Na\(^{+}\) ion is surrounded by six Cl\(^{-}\) ions, and every Cl\(^{-}\) ion is surrounded by six Na\(^{+}\) ions. Thus, there are many strong attractions between the positive and negative ions, which account for the high melting points of ionic compounds. For example, the melting point of NaCl is 801 °C. At room temperature, ionic compounds are solids.

![Sodium metal and Chlorine gas](image)

**Figure 6.1** The elements sodium and chlorine react to form the ionic compound sodium chloride, the compound that makes up table salt. The magnification of NaCl crystals shows the arrangement of Na\(^{+}\) and Cl\(^{-}\) ions.

**Q** What is the type of bonding between Na\(^{+}\) and Cl\(^{-}\) ions in NaCl?

Chemical Formulas of Ionic Compounds

The chemical formula of a compound represents the symbols and subscripts in the lowest whole-number ratio of the atoms or ions. In the formula of an ionic compound, the sum of the ionic charges in the formula is always zero. Thus, the total amount of positive charge is equal to the total amount of negative charge. For example, to achieve a stable electron configuration, one Na atom (metal) loses its one valence electron to form Na\(^{+}\), and one Cl atom (nonmetal) gains one electron to form a Cl\(^{-}\) ion. The formula NaCl indicates that the compound has charge balance because there is one sodium ion, Na\(^{+}\), for every chloride ion, Cl\(^{-}\).
ion, Cl\(^-\). Although the ions are positively or negatively charged, they are not shown in the formula of the compound.

### Subscripts in Formulas
Consider a compound of magnesium and chlorine. To achieve a stable electron configuration, one Mg atom (metal) loses its two valence electrons to form Mg\(^{2+}\). Two Cl atoms (nonmetals) each gain one electron to form two Cl\(^-\) ions. The two Cl\(^-\) ions are needed to balance the positive charge of Mg\(^{2+}\). This gives the formula MgCl\(_2\), magnesium chloride, in which the subscript 2 shows that two Cl\(^-\) ions are needed for charge balance.

### Writing Ionic Formulas from Ionic Charges
The subscripts in the formula of an ionic compound represent the number of positive and negative ions that give an overall charge of zero. Thus, we can now write a formula directly from the ionic charges of the positive and negative ions. Suppose we wish to write the formula for the ionic compound containing Na\(^+\) and S\(^{2-}\) ions. To balance the ionic charge of the S\(^{2-}\) ion, we will need to place two Na\(^+\) ions in the formula. This gives the formula Na\(_2\)S, which has an overall charge of zero. In the formula of an ionic compound, the cation is written first followed by the anion. Appropriate subscripts are used to show the number of each of the ions. This formula is the lowest ratio of ions in the ionic compound. Since ionic compounds do not exist as molecules, this lowest ratio of ions is called a *formula unit*.
### SAMPLE PROBLEM 6.3 Writing Formulas from Ionic Charges

Write the symbols for the ions, and the correct formula for the ionic compound formed when lithium and nitrogen react.

**TRY IT FIRST**

**SOLUTION**

Lithium, which is a metal in Group 1A (1), forms $\text{Li}^+$; nitrogen, which is a nonmetal in Group 5A (15), forms $\text{N}^{3–}$. The charge of $3–$ is balanced by three $\text{Li}^+$ ions.

$$3(1+) + 1(3–) = 0$$

Writing the cation (positive ion) first and the anion (negative ion) second gives the formula $\text{Li}_3\text{N}$.

**STUDY CHECK 6.3**

Write the symbols for the ions, and the correct formula for the ionic compound that would form when calcium and oxygen react.

**ANSWER**

$$\text{Ca}^{2+}, \text{O}^{2–}, \text{CaO}$$

### QUESTIONS AND PROBLEMS

#### 6.2 Ionic Compounds

**LEARNING GOAL** Using charge balance, write the correct formula for an ionic compound.

6.15 Which of the following pairs of elements are likely to form an ionic compound?

- a. lithium and chlorine
- b. oxygen and bromine
- c. potassium and oxygen
- d. sodium and neon
- e. cesium and magnesium
- f. nitrogen and fluorine

6.16 Which of the following pairs of elements are likely to form an ionic compound?

- a. helium and oxygen
- b. magnesium and chlorine
- c. chlorine and bromine
- d. potassium and sulfur
- e. sodium and potassium
- f. nitrogen and iodine

6.17 Write the correct ionic formula for the compound formed between each of the following pairs of ions:

- a. $\text{Na}^+$ and $\text{O}^{2–}$
- b. $\text{Al}^{3+}$ and $\text{Br}^–$
- c. $\text{Ba}^{2+}$ and $\text{N}^{3–}$
- d. $\text{Mg}^{2+}$ and $\text{F}^–$
- e. $\text{Al}^{3+}$ and $\text{S}^{2–}$

6.18 Write the correct ionic formula for the compound formed between each of the following pairs of ions:

- a. $\text{Al}^{3+}$ and $\text{Cl}^–$
- b. $\text{Ca}^{2+}$ and $\text{S}^{2–}$
- c. $\text{Li}^+$ and $\text{S}^{2–}$
- d. $\text{Rb}^+$ and $\text{P}^{3–}$
- e. $\text{Cs}^+$ and $\text{I}^–$

6.19 Write the symbols for the ions, and the correct formula for the ionic compound formed by each of the following:

- a. potassium and sulfur
- b. sodium and nitrogen
- c. aluminum and iodine
- d. gallium and oxygen

6.20 Write the symbols for the ions, and the correct formula for the ionic compound formed by each of the following:

- a. calcium and chlorine
- b. rubidium and bromine
- c. sodium and phosphorus
- d. magnesium and oxygen

### 6.3 Naming and Writing Ionic Formulas

**LEARNING GOAL** Given the formula of an ionic compound, write the correct name; given the name of an ionic compound, write the correct formula.

In the name of an ionic compound made up of two elements, the name of the metal ion, which is written first, is the same as its element name. The name of the nonmetal ion is obtained by using the first syllable of its element name followed by *ide*. In the name of any ionic compound, a space separates the name of the cation from the name of the anion. Subscripts are not used; they are understood because of the charge balance of the ions in the compound (see Table 6.4).
Iodized salt contains KI to prevent iodine deficiency.

### TABLE 6.4 Names of Some Ionic Compounds

<table>
<thead>
<tr>
<th>Compound</th>
<th>Metal Ion</th>
<th>Nonmetal Ion</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>KI</td>
<td>K⁺</td>
<td>I⁻</td>
<td>Potassium iodide</td>
</tr>
<tr>
<td>MgBr₂</td>
<td>Mg²⁺</td>
<td>Br⁻</td>
<td>Magnesium bromide</td>
</tr>
<tr>
<td>Al₂O₃</td>
<td>Al³⁺</td>
<td>O²⁻</td>
<td>Aluminum oxide</td>
</tr>
</tbody>
</table>

### SAMPLE PROBLEM 6.4 Naming Ionic Compounds

Write the name for the ionic compound Mg₃N₂.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg₃N₂</td>
<td>name</td>
<td>cation, anion</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Identify the cation and anion. The cation is Mg²⁺ and the anion is N³⁻.

**STEP 2** Name the cation by its element name. The cation Mg²⁺ is magnesium.

**STEP 3** Name the anion by using the first syllable of its element name followed by ide. The anion N³⁻ is nitride.

**STEP 4** Write the name for the cation first and the name for the anion second. magnesium nitride

**STUDY CHECK 6.4**

Name the compound Ga₂S₃.

**ANSWER**
gallium sulfide

### Metals with Variable Charges

We have seen that the charge of an ion of a representative element can be obtained from its group number. However, we cannot determine the charge of a transition element because it typically forms two or more positive ions. The transition elements lose electrons, but they are lost from the highest energy level and sometimes from a lower energy level as well. This is also true for metals of representative elements in Groups 4A (14) and 5A (15), such as Pb, Sn, and Bi.

In some ionic compounds, iron is in the Fe²⁺ form, but in other compounds, it has the Fe³⁺ form. Copper also forms two different ions, Cu⁺ and Cu²⁺. When a metal can form two or more types of ions, it has variable charge. Then we cannot predict the ionic charge from the group number.

For metals that form two or more ions, a naming system is used to identify the particular cation. To do this, a Roman numeral that is equal to the ionic charge is placed in parentheses immediately after the name of the metal. For example, Fe²⁺ is iron(II), and Fe³⁺ is iron(III). **TABLE 6.5** lists the ions of some metals that produce more than one ion.

**Why is a Roman numeral placed after the name of the cations of most transition elements?**
The transition elements form more than one positive ion except for zinc (Zn\(^{2+}\)), cadmium (Cd\(^{2+}\)), and silver (Ag\(^{+}\)), which form only one ion. Thus, no Roman numerals are used with zinc, cadmium, and silver when naming their cations in ionic compounds. Metals in Groups 4A (14) and 5A (15) also form more than one type of positive ion. For example, lead and tin in Group 4A (14) form cations with charges of 2+ and 4+, and bismuth in Group 5A (15) forms cations with charges of 3+ and 5+.

### Determination of Variable Charge

When you name an ionic compound, you need to determine if the metal is a representative element or a transition element. If it is a transition element, except for zinc, cadmium, or silver, you will need to use its ionic charge as a Roman numeral as part of its name. The calculation of ionic charge depends on the negative charge of the anions in the formula. For example, we use charge balance to determine the charge of a copper cation in the ionic compound \(\text{CuCl}_2\).

\[
\text{CuCl}_2 \quad \text{Cu charge} + 2 \text{Cl}^- \text{ charge} = 0
\]

\[
? + 2(1-) = 0
\]

\[
2+ + 2- = 0
\]

To indicate the 2+ charge for the copper ion \(\text{Cu}^{2+}\), we place the Roman numeral (II) immediately after copper when naming this compound: copper(II) chloride. Some ions and their location on the periodic table are seen in Figure 6.2.

### Table 6.5 Some Metals That Form More Than One Positive Ion

<table>
<thead>
<tr>
<th>Element</th>
<th>Possible Ions</th>
<th>Name of Ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bismuth</td>
<td>Bi(^{3+})</td>
<td>Bismuth(III)</td>
</tr>
<tr>
<td></td>
<td>Bi(^{5+})</td>
<td>Bismuth(V)</td>
</tr>
<tr>
<td>Chromium</td>
<td>Cr(^{2+})</td>
<td>Chromium(II)</td>
</tr>
<tr>
<td></td>
<td>Cr(^{3+})</td>
<td>Chromium(III)</td>
</tr>
<tr>
<td>Cobalt</td>
<td>Co(^{2+})</td>
<td>Cobalt(II)</td>
</tr>
<tr>
<td></td>
<td>Co(^{3+})</td>
<td>Cobalt(III)</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu(^{+})</td>
<td>Copper(I)</td>
</tr>
<tr>
<td></td>
<td>Cu(^{2+})</td>
<td>Copper(II)</td>
</tr>
<tr>
<td>Gold</td>
<td>Au(^{+})</td>
<td>Gold(I)</td>
</tr>
<tr>
<td></td>
<td>Au(^{3+})</td>
<td>Gold(II)</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe(^{2+})</td>
<td>Iron(II)</td>
</tr>
<tr>
<td></td>
<td>Fe(^{3+})</td>
<td>Iron(III)</td>
</tr>
<tr>
<td>Lead</td>
<td>Pb(^{2+})</td>
<td>Lead(II)</td>
</tr>
<tr>
<td></td>
<td>Pb(^{4+})</td>
<td>Lead(IV)</td>
</tr>
<tr>
<td>Manganese</td>
<td>Mn(^{2+})</td>
<td>Manganese(II)</td>
</tr>
<tr>
<td></td>
<td>Mn(^{3+})</td>
<td>Manganese(III)</td>
</tr>
<tr>
<td>Mercury</td>
<td>Hg(^{2+})</td>
<td>Mercury(I)*</td>
</tr>
<tr>
<td></td>
<td>Hg(^{2+})</td>
<td>Mercury(II)</td>
</tr>
<tr>
<td>Nickel</td>
<td>Ni(^{2+})</td>
<td>Nickel(II)</td>
</tr>
<tr>
<td></td>
<td>Ni(^{3+})</td>
<td>Nickel(III)</td>
</tr>
<tr>
<td>Tin</td>
<td>Sn(^{2+})</td>
<td>Tin(II)</td>
</tr>
<tr>
<td></td>
<td>Sn(^{4+})</td>
<td>Tin(IV)</td>
</tr>
</tbody>
</table>

*Mercury(II) ions form an ion pair with a 2+ charge.
6.3 Naming and Writing Ionic Formulas

SAMPLE PROBLEM 6.5 Naming Ionic Compounds with Variable Charge Metal Ions

Antifouling paint contains Cu₂O, which prevents the growth of barnacles and algae on the bottoms of boats. What is the name of Cu₂O?

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Cu₂O</td>
<td>name</td>
<td>cation, anion</td>
</tr>
</tbody>
</table>

STEP 1 Determine the charge of the cation from the anion.

<table>
<thead>
<tr>
<th>Element</th>
<th>Metal</th>
<th>Nonmetal</th>
</tr>
</thead>
<tbody>
<tr>
<td>copper (Cu)</td>
<td>oxygen (O)</td>
<td></td>
</tr>
<tr>
<td>Group</td>
<td>transition element</td>
<td>6A (16)</td>
</tr>
<tr>
<td>Ion</td>
<td>Cu²⁺</td>
<td>O²⁻</td>
</tr>
<tr>
<td>Charge Balance</td>
<td>( \frac{2\text{Cu}^2\text{Br}^\text{Br}}{2} = \frac{2+}{2} = 1+ )</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Name the cation by its element name and use a Roman numeral in parentheses for the charge. copper(I)

STEP 3 Name the anion by using the first syllable of its element name followed by ide. oxide

STEP 4 Write the name for the cation first and the name for the anion second. copper(I) oxide

STUDY CHECK 6.5

Write the name for the compound with the formula Mn₂S₃.

ANSWER

manganese(III) sulfide

Guide to Naming Ionic Compounds with Variable Charge Metals

STEP 1 Determine the charge of the cation from the anion.

STEP 2 Name the cation by its element name and use a Roman numeral in parentheses for the charge.

STEP 3 Name the anion by using the first syllable of its element name followed by ide.

STEP 4 Write the name for the cation first and the name for the anion second.

TABLE 6.6 Some Ionic Compounds of Metals That Form Two Kinds of Positive Ions

<table>
<thead>
<tr>
<th>Compound</th>
<th>Systematic Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>FeCl₂</td>
<td>Iron(II) chloride</td>
</tr>
<tr>
<td>Fe₂O₃</td>
<td>Iron(III) oxide</td>
</tr>
<tr>
<td>Cu₃P</td>
<td>Copper(I) phosphide</td>
</tr>
<tr>
<td>CrBr₂</td>
<td>Chromium(II) bromide</td>
</tr>
<tr>
<td>SnCl₂</td>
<td>Tin(II) chloride</td>
</tr>
<tr>
<td>PbS₂</td>
<td>Lead(IV) sulfide</td>
</tr>
<tr>
<td>BiF₃</td>
<td>Bismuth(III) fluoride</td>
</tr>
</tbody>
</table>

Writing Formulas from the Name of an Ionic Compound

The formula for an ionic compound is written from the first part of the name that describes the metal ion, including its charge, and the second part of the name that specifies the nonmetal ion. Subscripts are added, as needed, to balance the charge. The steps for writing a formula from the name of an ionic compound are shown in Sample Problem 6.6.
Guide to Writing Formulas from the Name of an Ionic Compound

**STEP 1** Identify the cation and anion.

**STEP 2** Balance the charges.

**STEP 3** Write the formula, cation first, using subscripts from the charge balance.

---

**SAMPLE PROBLEM 6.6 Writing Formulas for Ionic Compounds**

Write the correct formula for iron(III) chloride.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>iron(III) chloride</td>
<td>formula</td>
<td>cation, anion, charge balance</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Identify the cation and anion.

<table>
<thead>
<tr>
<th>Type of Ion</th>
<th>Cation</th>
<th>Anion</th>
</tr>
</thead>
<tbody>
<tr>
<td>iron(III)</td>
<td>Fe⁺³</td>
<td>Cl⁻</td>
</tr>
<tr>
<td>Group</td>
<td>transition element</td>
<td>7A (17)</td>
</tr>
<tr>
<td>Symbol of Ion</td>
<td>Fe⁺³</td>
<td>Cl⁻</td>
</tr>
</tbody>
</table>

**STEP 2** Balance the charges. The charge of 3+ is balanced by three Cl⁻ ions.

\[ 1(3^+) + 3(1^-) = 0 \]

**STEP 3** Write the formula, cation first, using subscripts from the charge balance.

\[ \text{FeCl}_3 \]

**STUDY CHECK 6.6**

Write the correct formula for chromium(III) oxide.

**ANSWER**

\[ \text{Cr}_2\text{O}_3 \]

---

**QUESTIONS AND PROBLEMS**

**6.3 Naming and Writing Ionic Formulas**

**LEARNING GOAL** Given the formula of an ionic compound, write the correct name; given the name of an ionic compound, write the correct formula.

**6.21** Write the name for each of the following ionic compounds:

\[ \text{a. Al}_2\text{O}_3 \]  \[ \text{b. CaCl}_2 \]  \[ \text{c. Na}_2\text{O} \]
\[ \text{d. Mg}_3\text{P}_2 \]  \[ \text{e. } \text{K} \text{I} \]  \[ \text{f. BaF}_2 \]

**6.22** Write the name for each of the following ionic compounds:

\[ \text{a. MgCl}_2 \]  \[ \text{b. } \text{K}_2\text{P} \]  \[ \text{c. Li}_2\text{S} \]
\[ \text{d. CaF} \]  \[ \text{e. MgO} \]  \[ \text{f. SrBr}_2 \]

**6.23** Write the name for each of the following ions (include the Roman numeral when necessary):

\[ \text{a. Fe}^{2+} \]  \[ \text{b. Cu}^{2+} \]  \[ \text{c. Zn}^{2+} \]
\[ \text{d. Pb}^{2+} \]  \[ \text{e. Cr}^{3+} \]  \[ \text{f. Mn}^{2+} \]

**6.24** Write the name for each of the following ions (include the Roman numeral when necessary):

\[ \text{a. Ag}^+ \]  \[ \text{b. Cu}^+ \]  \[ \text{c. Bi}^{3+} \]
\[ \text{d. Sn}^{2+} \]  \[ \text{e. Au}^{3+} \]  \[ \text{f. Ni}^{2+} \]

**6.25** Write the name for each of the following ionic compounds:

\[ \text{a. SnCl}_2 \]  \[ \text{b. FeO} \]  \[ \text{c. Cu}_2\text{S} \]
\[ \text{d. CuS} \]  \[ \text{e. CdBr}_2 \]  \[ \text{f. HgCl}_2 \]

**6.26** Write the name for each of the following ionic compounds:

\[ \text{a. Ag}_2\text{P} \]  \[ \text{b. PbS} \]  \[ \text{c. SnO}_2 \]
\[ \text{d. MnCl}_3 \]  \[ \text{e. Bi}_2\text{O}_3 \]  \[ \text{f. CoCl}_2 \]

**6.27** Write the symbol for the cation in each of the following ionic compounds:

\[ \text{a. AuCl}_3 \]  \[ \text{b. Fe}_2\text{O}_3 \]  \[ \text{c. Pb}_4 \]  \[ \text{d. SnCl}_2 \]

**6.28** Write the symbol for the cation in each of the following ionic compounds:

\[ \text{a. FeCl}_2 \]  \[ \text{b. CrO} \]  \[ \text{c. Ni}_2\text{S}_3 \]  \[ \text{d. AlP} \]

**6.29** Write the formula for each of the following ionic compounds:

\[ \text{a. } \text{magnesium chloride} \]  \[ \text{b. } \text{sodium sulfide} \]
\[ \text{c. } \text{copper(I) oxide} \]  \[ \text{d. } \text{zinc phosphide} \]
\[ \text{e. } \text{gold(III) nitride} \]  \[ \text{f. } \text{cobalt(III) fluoride} \]
6.4 Polyatomic Ions

**LEARNING GOAL** Write the name and formula for an ionic compound containing a polyatomic ion.

An ionic compound may also contain a *polyatomic ion* as one of its cations or anions. A *polyatomic ion* is a group of covalently bonded atoms that has an overall ionic charge. Most polyatomic ions consist of a nonmetal such as phosphorus, sulfur, carbon, or nitrogen covalently bonded to oxygen atoms.

Almost all the polyatomic ions are anions with charges 1\(^-\), 2\(^-\), or 3\(^-\). Only one common polyatomic ion, \(\text{NH}_4^+\), has a positive charge. Some models of common polyatomic ions are shown in **FIGURE 6.3**.

**Names of Polyatomic Ions**

The names of the most common polyatomic ions end in *ate*, such as nitrate and sulfate. When a related ion has one less oxygen atom, the *ite* ending is used for its name such as nitrite and sulfite. Recognizing these endings will help you identify polyatomic ions in the name of a compound. The hydroxide ion (\(\text{OH}^-\)) and cyanide ion (\(\text{CN}^-\)) are exceptions to this naming pattern.

**FIGURE 6.3** Many products contain polyatomic ions, which are groups of atoms that have an ionic charge.

Q What is the charge of a sulfate ion?
By learning the formulas, charges, and names of the polyatomic ions shown in bold type in Table 6.7, you can derive the related ions. Note that both the \textit{ate} ion and \textit{ite} ion of a particular nonmetal have the same ionic charge. For example, the sulfate ion is \( \text{SO}_4^{2-} \), and the sulfite ion, which has one less oxygen atom, is \( \text{SO}_3^{2-} \). Phosphate and phosphite ions each have a 3⁻ charge; nitrate and nitrite each have a 1⁻ charge; and perchlorate, chlorate, chlorite, and hypochlorite all have a 1⁻ charge. The halogens form four different polyatomic ions with oxygen.

### Table 6.7: Names and Formulas of Some Common Polyatomic Ions

<table>
<thead>
<tr>
<th>Nonmetal</th>
<th>Formula of Ion*</th>
<th>Name of Ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>( \text{OH}^- )</td>
<td>Hydroxide</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>( \text{NH}_4^+ )</td>
<td>Ammonium</td>
</tr>
<tr>
<td></td>
<td>( \text{NO}_3^- )</td>
<td>Nitrate</td>
</tr>
<tr>
<td></td>
<td>( \text{NO}_2^- )</td>
<td>Nitrite</td>
</tr>
<tr>
<td>Chlorine</td>
<td>( \text{ClO}_4^- )</td>
<td>Perchlorate</td>
</tr>
<tr>
<td></td>
<td>( \text{ClO}_3^- )</td>
<td>Chlorate</td>
</tr>
<tr>
<td></td>
<td>( \text{ClO}_2^- )</td>
<td>Chlorite</td>
</tr>
<tr>
<td></td>
<td>( \text{ClO}^- )</td>
<td>Hypochlorite</td>
</tr>
<tr>
<td>Carbon</td>
<td>( \text{CO}_2^{2-} )</td>
<td>Carbonate</td>
</tr>
<tr>
<td></td>
<td>( \text{HCO}_3^- )</td>
<td>Hydrogen carbonate (or bicarbonate)</td>
</tr>
<tr>
<td></td>
<td>( \text{CN}^- )</td>
<td>Cyanide</td>
</tr>
<tr>
<td></td>
<td>( \text{C}_2\text{H}_3\text{O}_2^- )</td>
<td>Acetate</td>
</tr>
<tr>
<td>Sulfur</td>
<td>( \text{SO}_2^{2-} )</td>
<td>Sulfate</td>
</tr>
<tr>
<td></td>
<td>( \text{HSO}_4^- )</td>
<td>Hydrogen sulfate (or bisulfate)</td>
</tr>
<tr>
<td></td>
<td>( \text{SO}_3^{2-} )</td>
<td>Sulfite</td>
</tr>
<tr>
<td></td>
<td>( \text{HSO}_3^- )</td>
<td>Hydrogen sulfite (or bisulfite)</td>
</tr>
<tr>
<td>Phosphorus</td>
<td>( \text{PO}_4^{3-} )</td>
<td>Phosphate</td>
</tr>
<tr>
<td></td>
<td>( \text{HPO}_4^{2-} )</td>
<td>Hydrogen phosphate</td>
</tr>
<tr>
<td></td>
<td>( \text{H}_2\text{PO}_4^- )</td>
<td>Dihydrogen phosphate</td>
</tr>
<tr>
<td></td>
<td>( \text{PO}_3^{3-} )</td>
<td>Phosphite</td>
</tr>
</tbody>
</table>

*Formulas and names in bold type indicate the most common polyatomic ion for that element.

The formula of hydrogen carbonate, or \textit{bicarbonate}, is written by placing a hydrogen in front of the polyatomic ion formula for carbonate (\( \text{CO}_3^{2-} \)), and the charge is decreased from 2⁻ to 1⁻ to give \( \text{HCO}_3^- \).

\[ \text{H}^+ + \text{CO}_3^{2-} = \text{HCO}_3^- \]

### Writing Formulas for Compounds Containing Polyatomic Ions

No polyatomic ion exists by itself. Like any ion, a polyatomic ion must be associated with ions of opposite charge. The bonding between polyatomic ions and other ions is one of electrical attraction. For example, the compound sodium chlorite consists of sodium ions (\( \text{Na}^+ \)) and chlorite ions (\( \text{ClO}_3^- \)) held together by ionic bonds.

To write correct formulas for compounds containing polyatomic ions, we follow the same rules of charge balance that we used for writing the formulas for simple ionic compounds. The total negative and positive charges must equal zero. For example, consider the formula for a compound containing sodium ions and chlorite ions. The ions are written as

\[ \text{Na}^+ \text{ClO}_3^- \]

Sodium ion    Chlorite ion
\((1^+) + (1^-) = 0\)

Because one ion of each balances the charge, the formula is written as

\[ \text{NaClO}_3 \]

Sodium chlorite
When more than one polyatomic ion is needed for charge balance, parentheses are used to enclose the formula of the ion. A subscript is written outside the right parenthesis of the polyatomic ion to indicate the number needed for charge balance. Consider the formula for magnesium nitrate. The ions in this compound are the magnesium ion and the nitrate ion, a polyatomic ion.

$$\text{Mg}^{2+} \quad \text{NO}_3^-$$

Magnesium ion  Nitrate ion

To balance the positive charge of $2^+$ on the magnesium ion, two nitrate ions are needed. In the formula of the compound, parentheses are placed around the nitrate ion, and the subscript 2 is written outside the right parenthesis.

$$\text{Mg}^2+ \quad \text{NO}_3^- \quad \text{NO}_3^- \quad \text{Mg(NO}_3)_2$$

Parentheses enclose the formula of the nitrate ion  Subscript outside the parenthesis indicates the use of two nitrate ions

**SAMPLE PROBLEM 6.7** Writing Formulas Containing Polyatomic Ions

An antacid called Amphojel contains aluminum hydroxide, which treats acid indigestion and heartburn. Write the formula for aluminum hydroxide.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminium hydroxide</td>
<td>formula</td>
<td>cation, polyatomic anion</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Identify the cation and polyatomic ion (anion).

- Cation: aluminum
- Polyatomic ion (anion): hydroxide
- Aluminium hydroxide

**STEP 2** Balance the charges. The charge of $3^+$ is balanced by three $\text{OH}^-$ ions.

$$1(3+) + 3(1-) = 0$$

**STEP 3** Write the formula, cation first, using the subscripts from charge balance. The formula for the compound is written by enclosing the formula of the hydroxide ion, $\text{OH}^-$, in parentheses and writing the subscript 3 outside the right parenthesis.

$$\text{Al(OH)}_3$$

**STUDY CHECK 6.7**

Write the formula for a compound containing ammonium ions and phosphate ions.

**ANSWER**

$$(\text{NH}_4)_3\text{PO}_4$$
Naming Ionic Compounds Containing Polyatomic Ions

When naming ionic compounds containing polyatomic ions, we first write the positive ion, usually a metal, and then we write the name for the polyatomic ion. It is important that you learn to recognize the polyatomic ion in the formula and name it correctly. As with other ionic compounds, no prefixes are used.

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
<th>Medical Use</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na₂SO₄</td>
<td>Sodium sulfate</td>
<td></td>
</tr>
<tr>
<td>FePO₄</td>
<td>Iron(III) phosphate</td>
<td></td>
</tr>
<tr>
<td>Al₂(CO₃)₃</td>
<td>Aluminum carbonate</td>
<td></td>
</tr>
</tbody>
</table>

TABLE 6.8 lists the formulas and names of some ionic compounds that include polyatomic ions and also gives their uses in medicine and industry.

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
<th>Medical Use</th>
</tr>
</thead>
<tbody>
<tr>
<td>AlPO₄</td>
<td>Aluminum phosphate</td>
<td>Antacid</td>
</tr>
<tr>
<td>Al₂(SO₄)₃</td>
<td>Aluminum sulfate</td>
<td>Antiperspirant, anti-infective</td>
</tr>
<tr>
<td>BaSO₄</td>
<td>Barium sulfate</td>
<td>Contrast medium for X-rays</td>
</tr>
<tr>
<td>CaCO₃</td>
<td>Calcium carbonate</td>
<td>Antacid, calcium supplement</td>
</tr>
<tr>
<td>Ca₃(PO₄)₂</td>
<td>Calcium phosphate</td>
<td>Calcium dietary supplement</td>
</tr>
<tr>
<td>CaSO₄</td>
<td>Calcium sulfate</td>
<td>Plaster casts</td>
</tr>
<tr>
<td>MgSO₄</td>
<td>Magnesium sulfate</td>
<td>Cathartic, Epsom salts</td>
</tr>
<tr>
<td>K₂CO₃</td>
<td>Potassium carbonate</td>
<td>Alkalizer, diuretic</td>
</tr>
<tr>
<td>AgNO₃</td>
<td>Silver nitrate</td>
<td>Topical anti-infective</td>
</tr>
<tr>
<td>NaHCO₃</td>
<td>Sodium bicarbonate or sodium hydrogen carbonate</td>
<td>Antacid</td>
</tr>
<tr>
<td>Zn₃(PO₄)₂</td>
<td>Zinc phosphate</td>
<td>Dental cement</td>
</tr>
</tbody>
</table>

A solution of Epsom salts, magnesium sulfate, MgSO₄, may be used to soothe sore muscles.

SAMPLE PROBLEM 6.8 Naming Compounds Containing Polyatomic Ions

Name the following ionic compounds:

a. Cu(NO₂)₂
b. KClO₃

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>formula</td>
<td>name</td>
<td>cation, polyatomic ion</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Formula</th>
<th>Cation</th>
<th>Anion</th>
<th>Name of Cation</th>
<th>Name of Anion</th>
<th>Name of Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. Cu(NO₂)₂</td>
<td>Cu²⁺</td>
<td>NO₂⁻</td>
<td>Copper(II) ion</td>
<td>Nitrite ion</td>
<td>Copper(II) nitrite</td>
</tr>
<tr>
<td>b. KClO₃</td>
<td>K⁺</td>
<td>ClO₃⁻</td>
<td>Potassium ion</td>
<td>Chlorate ion</td>
<td>Potassium chlorate</td>
</tr>
</tbody>
</table>

STUDY CHECK 6.8

What is the name of Co₃(PO₄)₂?

ANSWER

cobalt(II) phosphate
**6.5 Molecular Compounds: Sharing Electrons**

**LEARNING GOAL** Given the formula of a molecular compound, write its correct name; given the name of a molecular compound, write its formula.

A molecular compound consists of atoms of two or more nonmetals that share one or more valence electrons. The shared atoms are held together by covalent bonds that form a molecule. There are many more molecular compounds than there are ionic ones. For example, water (H₂O) and carbon dioxide (CO₂) are both molecular compounds. Molecular compounds consist of molecules, which are discrete groups of atoms in a definite proportion. A molecule of water (H₂O) consists of two atoms of hydrogen and one atom of oxygen. When you have iced tea, perhaps you add molecules of sugar (C₁₂H₂₂O₁₁), which is a molecular compound. Other familiar molecular compounds include propane (C₃H₈), alcohol (C₂H₅OH), the antibiotic amoxicillin (C₁₆H₁₈N₅O₈S), and the antidepressant Prozac (C₁₇H₁₈F₃NO).

**Names and Formulas of Molecular Compounds**

When naming a molecular compound, the first nonmetal in the formula is named by its element name; the second nonmetal is named using the first syllable of its element name,
ChApter 6
Ionic and Molecular Compounds

followed by ide. When a subscript indicates two or more atoms of an element, a prefix is shown in front of its name. TABLE 6.9 lists prefixes used in naming molecular compounds.

The names of molecular compounds need prefixes because several different compounds can be formed from the same two nonmetals. For example, carbon and oxygen can form two different compounds, carbon monoxide, CO, and carbon dioxide, CO2, in which the number of atoms of oxygen in each compound is indicated by the prefixes mono or di in their names.

When the vowels o and o or a and o appear together, the first vowel is omitted, as in carbon monoxide. In the name of a molecular compound, the prefix mono is usually omitted, as in NO, nitrogen oxide. Traditionally, however, CO is named carbon monoxide. TABLE 6.10 lists the formulas, names, and commercial uses of some molecular compounds.

| TABLE 6.9 Prefixes Used in Naming Molecular Compounds |
|-----------------|-----|-----|
| 1 mono          | 6   | hexa |
| 2 di            | 7   | hept |
| 3 tri           | 8   | octa |
| 4 tetra         | 9   | nona |
| 5 penta         | 10  | deca |

<p>| TABLE 6.10 Some Common Molecular Compounds |
|-------------------------------|----------------|----------------|</p>
<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
<th>Commercial Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>CO2</td>
<td>Carbon dioxide</td>
<td>Fire extinguishers, dry ice, propellant in aerosols,</td>
</tr>
<tr>
<td></td>
<td></td>
<td>carbonation of beverages</td>
</tr>
<tr>
<td>CS2</td>
<td>Carbon disulfide</td>
<td>Manufacture of rayon</td>
</tr>
<tr>
<td>N2O</td>
<td>Dinitrogen oxide</td>
<td>Inhalation anesthetic, “laughing gas”</td>
</tr>
<tr>
<td>NO</td>
<td>Nitrogen oxide</td>
<td>Stabilizer, biochemical messenger in cells</td>
</tr>
<tr>
<td>SO2</td>
<td>Sulfur dioxide</td>
<td>Preserving fruits, vegetables; disinfectant in breweries; bleaching textiles</td>
</tr>
<tr>
<td>SF6</td>
<td>Sulfur hexafluoride</td>
<td>Electrical circuits</td>
</tr>
<tr>
<td>SO3</td>
<td>Sulfur trioxide</td>
<td>Manufacture of explosives</td>
</tr>
</tbody>
</table>

Why are prefixes used to name molecular compounds?

SAMPLE PROBLEM 6.9 Naming Molecular Compounds

Name the molecular compound NCl3.

Try it First

Solution

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>NCl3</td>
<td>name</td>
<td>prefixes</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Name the first nonmetal by its element name. In NCl3, the first nonmetal (N) is nitrogen.

**STEP 2** Name the second nonmetal by using the first syllable of its element name followed by ide. The second nonmetal (Cl) is named chloride.

**STEP 3** Add prefixes to indicate the number of atoms (subscripts). Because there is one nitrogen atom, no prefix is needed. The subscript 3 for the Cl atoms is shown as the prefix tri.

<table>
<thead>
<tr>
<th>Symbol of Element</th>
<th>N</th>
<th>Cl</th>
</tr>
</thead>
<tbody>
<tr>
<td>Name</td>
<td>nitrogen</td>
<td>chloride</td>
</tr>
<tr>
<td>Subscript</td>
<td>1</td>
<td>3</td>
</tr>
<tr>
<td>Prefix</td>
<td>none</td>
<td>tri</td>
</tr>
</tbody>
</table>

The name of NCl3 is nitrogen trichloride.

STuDY CHeCK 6.9
Write the name for each of the following molecular compounds:

a. SiBr4

b. Br2O

ANSwer

a. silicon tetrabromide

b. dibromine oxide
Writing Formulas from the Names of Molecular Compounds

In the name of a molecular compound, the names of two nonmetals are given along with prefixes for the number of atoms of each. To write the formula from the name, we use the symbol for each element and a subscript if a prefix indicates two or more atoms.

SAMPLE PROBLEM 6.10 Writing Formulas for Molecular Compounds

Write the formula for the molecular compound diboron trioxide.

TRY IT FIRST

SOLUTION

STEP 1 Write the symbols in the order of the elements in the name.

```
<table>
<thead>
<tr>
<th>Name of Element</th>
<th>boron</th>
<th>oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>Symbol of Element</td>
<td>B</td>
<td>O</td>
</tr>
<tr>
<td>Subscript</td>
<td>2 (di)</td>
<td>3 (tri)</td>
</tr>
</tbody>
</table>
```

STEP 2 Write any prefixes as subscripts. The prefix *di* in *diboron* indicates that there are two atoms of boron, shown as a subscript 2 in the formula. The prefix *tri* in *trioxide* indicates that there are three atoms of oxygen, shown as a subscript 3 in the formula.

\[ \text{B}_2\text{O}_3 \]

STUDY CHECK 6.10

Write the formula for the molecular compound iodine pentafluoride.

ANSWER

IF₅

Summary of Naming Ionic and Molecular Compounds

We have now examined strategies for naming ionic and molecular compounds. In general, compounds having two elements are named by stating the first element name followed by the name of the second element with an *ide* ending. If the first element is a metal, the compound is usually ionic; if the first element is a nonmetal, the compound is usually molecular. For ionic compounds, it is necessary to determine whether the metal can form more than one type of positive ion; if so, a Roman numeral following the name of the metal indicates the particular ionic charge. One exception is the ammonium ion, \( \text{NH}_4^+ \), which is also written first as a positively charged polyatomic ion. Ionic compounds having three or more elements include some type of polyatomic ion. They are named by ionic rules but have an *ate* or *ite* ending when the polyatomic ion has a negative charge.

In naming molecular compounds having two elements, prefixes are necessary to indicate two or more atoms of each nonmetal as shown in that particular formula (see FIGURE 6.4).

SAMPLE PROBLEM 6.11 Naming Ionic and Molecular Compounds

Identify each of the following compounds as ionic or molecular and give its name:

a. K₃P 
   b. NiSO₄ 
   c. SO₃

TRY IT FIRST

SOLUTION

a. K₃P, consisting of a metal and a nonmetal, is an ionic compound. As a representative element in Group 1A (1), K forms the potassium ion, \( \text{K}^+ \). Phosphorus, as a representative element in Group 5A (15), forms a phosphide ion, \( \text{P}^{3-} \). Writing the name of the cation followed by the name of the anion gives the name potassium phosphide.

b. NiSO₄

SOLUTION

b. NiSO₄
b. NiSO₄, consisting of a cation of a transition element and a polyatomic ion \( \text{SO}_4^{2-} \), is an ionic compound. As a transition element, Ni forms more than one type of ion. In this formula, the \( 2^- \) charge of \( \text{SO}_4^{2-} \) is balanced by one nickel ion, \( \text{Ni}^{2+} \). In the name, a Roman numeral written after the metal name, nickel(II), specifies the \( 2^+ \) charge. The anion \( \text{SO}_4^{2-} \) is a polyatomic ion named sulfate. The compound is named nickel(II) sulfate.

c. \( \text{SO}_3 \) consists of two nonmetals, which indicates that it is a molecular compound. The first element S is sulfur (no prefix is needed). The second element O, oxide, has subscript 3, which requires a prefix \( tri \) in the name. The compound is named sulfur trioxide.

**STUDY CHECK 6.11**

What is the name of \( \text{Fe(NO}_3)_3 \)?

**ANSWER**

iron(III) nitrate
6.52 Write the formula for each of the following molecular compounds:
   a. sulfur dioxide  
   b. silicon tetrachloride  
   c. iodine trifluoride  
   d. dinitrogen oxide

6.53 Write the formula for each of the following molecular compounds:
   a. oxygen difluoride  
   b. boron trichloride  
   c. dinitrogen trioxide  
   d. sulfur hexafluoride

6.54 Write the formula for each of the following molecular compounds:
   a. sulfur dibromide  
   b. carbon disulfide  
   c. tetraphosphorus hexoxide  
   d. dinitrogen pentoxide

Applications

6.55 Name each of the following ionic or molecular compounds:
   a. Al₂(SO₄)₃, antiperspirant  
   b. CaCO₃, antacid  
   c. N₂O, “laughing gas,” inhaled anesthetic  
   d. Mg(OH)₂, laxative

6.56 Name each of the following ionic or molecular compounds:
   a. Al₂(SO₄)₃, antacid  
   b. FeSO₄, iron supplement in vitamins  
   c. NO, vasodilator  
   d. Cu(OH)₂, fungicide

Follow Up

COMPOUNDS AT THE PHARMACY

A few days ago, Richard went back to the pharmacy to pick up aspirin, C₉H₈O₄, and acetaminophen, C₈H₉NO₂. He also wanted to talk to Sarah about a way to treat his sore toe. Sarah recommended soaking his foot in a solution of Epsom salts, which is magnesium sulfate. Richard also asked Sarah to recommend an antacid for his upset stomach and an iron supplement. Sarah suggested an antacid that contains calcium carbonate and aluminum hydroxide, and iron(II) sulfate as an iron supplement. Richard also picked up toothpaste containing tin(II) fluoride, and carbonated water, which contains carbon dioxide.

Applications

6.57 Write the chemical formula for each of the following:
   a. magnesium sulfate  
   b. tin(II) fluoride  
   c. aluminum hydroxide

6.58 Write the chemical formula for each of the following:
   a. calcium carbonate  
   b. carbon dioxide  
   c. iron(II) sulfate

6.59 Identify each of the compounds in problem 6.57 as ionic or molecular.

6.60 Identify each of the compounds in problem 6.58 as ionic or molecular.
6.1 Ions: Transfer of Electrons

LEARNING GOAL Write the symbols for the simple ions of the representative elements.

- The stability of the noble gases is associated with a mono... the name of an ionic compound.  
- In naming ionic compounds, the positive ion is given first followed by the name of the negative ion.
- The names of ionic compounds containing two elements end with ide.
- Except for Ag, Cd, and Zn, transition elements form cations with two or more ionic charges.
- The charge of the cation is determined from the total negative charge in the formula and included as a Roman numeral immediately following the name of the metal that has a variable charge.

6.4 Polyatomic Ions

LEARNING GOAL Write the name and formula for an ionic compound containing a polyatomic ion.

- A polyatomic ion is a covalently bonded group of atoms with an electrical charge; for example, the carbonate ion has the formula $\text{CO}_3^{2-}$.
- Most polyatomic ions have names that end with ate or ite.
- Most polyatomic ions contain a nonmetal and one or more oxygen atoms.
- The ammonium ion, $\text{NH}_4^+$, is a positive polyatomic ion.
- When more than one polyatomic ion is used for charge balance, parentheses enclose the formula of the polyatomic ion.

6.5 Molecular Compounds: Sharing Electrons

LEARNING GOAL Given the formula of a molecular compound, write its correct name; given the name of a molecular compound, write its formula.

- In a covalent bond, atoms of nonmetals share valence electrons such that each atom has a stable electron configuration.
- The first nonmetal in a molecular compound uses its element name; the second nonmetal uses the first syllable of its element name followed by ide.
- The name of a molecular compound with two different atoms uses prefixes to indicate the subscripts in the formula.

<table>
<thead>
<tr>
<th>Number</th>
<th>Prefix</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>mono</td>
</tr>
<tr>
<td>2</td>
<td>di</td>
</tr>
<tr>
<td>3</td>
<td>tri</td>
</tr>
<tr>
<td>4</td>
<td>tetra</td>
</tr>
<tr>
<td>5</td>
<td>penta</td>
</tr>
</tbody>
</table>

KEY TERMS

- **anion** A negatively charged ion such as $\text{Cl}^-$, $\text{O}_2^-$, or $\text{SO}_4^{2-}$.
- **cation** A positively charged ion such as $\text{Na}^+$, $\text{Mg}^{2+}$, $\text{Al}^{3+}$, or $\text{NH}_4^+$.
- **chemical formula** The group of symbols and subscripts that represents the atoms or ions in a compound.
- **covalent bond** A sharing of valence electrons by atoms.
- **ion** An atom or group of atoms having an electrical charge because of a loss or gain of electrons.
- **ionic charge** The difference between the number of protons (positive) and the number of electrons (negative) written in the upper right corner of the symbol for the element or polyatomic ion.
- **ionic compound** A compound of positive and negative ions held together by ionic bonds.
- **molecular compound** A combination of atoms in which stable electron configurations are attained by sharing electrons.
- **molecule** The smallest unit of two or more atoms held together by covalent bonds.
- **octet** A set of eight valence electrons.
- **octet rule** Elements in Groups 1A to 7A (1, 2, 13 to 17) react with other elements by forming ionic or covalent bonds to produce a stable electron configuration, usually eight electrons in the outer shell.
- **polyatomic ion** A group of covalently bonded nonmetal atoms that has an overall electrical charge.
Writing Positive and Negative Ions (6.1)

- In the formation of an ionic bond, atoms of a metal lose and atoms of a nonmetal gain valence electrons to acquire a stable electron configuration, usually eight valence electrons.
- This tendency of atoms to attain a stable electron configuration is known as the octet rule.

Example: State the number of electrons lost or gained by atoms and the ion formed for each of the following to obtain a stable electron configuration:

- a. Br
- b. Ca
- c. S

Answer: a. Br atoms gain one electron to achieve a stable electron configuration, Br⁻.
- b. Ca atoms lose two electrons to achieve a stable electron configuration, Ca²⁺.
- c. S atoms gain two electrons to achieve a stable electron configuration, S²⁻.

Writing Ionic Formulas (6.2)

- The chemical formula of a compound represents the lowest whole-number ratio of the atoms or ions.
- In the chemical formula of an ionic compound, the sum of the positive and negative charges is always zero.
- Thus, in a chemical formula of an ionic compound, the total positive charge is equal to the total negative charge.

Example: Write the formula for magnesium phosphide.

Answer: Magnesium phosphide is an ionic compound that contains the ions Mg²⁺ and P³⁻.
Using charge balance, we determine the number(s) of each type of ion.
  3(2+) + 2(3-) = 0
  3Mg²⁺ and 2P³⁻ give the formula Mg₃P₂.

Naming Ionic Compounds (6.3)

- In the name of an ionic compound made up of two elements, the name of the metal ion, which is written first, is the same as its element name.
- For metals that form two or more ions, a Roman numeral that is equal to the ionic charge is placed in parentheses immediately after the name of the metal.
- The name of a nonmetal ion is obtained by using the first syllable of its element name followed by ide.

Example: What is the name of PbS?

Answer: This compound contains the S²⁻ ion which has a 2⁻ charge.
For charge balance, the positive ion must have a charge of 2⁺.
  Pb° + (2−) = 0; Pb = 2⁺
Because lead can form two different positive ions, a Roman numeral (II) is used in the name of the compound: lead(II) sulfide.

Writing the Names and Formulas for Molecular Compounds (6.5)

- When naming a molecular compound, the first nonmetal in the formula is named by its element name; the second nonmetal is named using the first syllable of its element name followed by ide.
- When a subscript indicates two or more atoms of an element, a prefix is shown in front of its name.

Example: Name the molecular compound BrF₅.

Answer: Two nonmetals share electrons and form a molecular compound. Br (first nonmetal) is bromine; F (second nonmetal) is fluoride. In the name for a molecular compound, prefixes indicate the subscripts in the formulas. The subscript 1 is understood for Br. The subscript 5 for fluoride is written with the prefix penta. The name is bromine pentafluoride.
6.65 Consider the following Lewis symbols for elements X and Y:

\[ \text{X} \quad \text{Y} \]

a. What are the group numbers of X and Y?
b. Will a compound of X and Y be ionic or molecular?
c. What ions would be formed by X and Y?
d. What would be the formula of a compound of X and Y?
e. What would be the formula of a compound of X and sulfur?
f. What would be the formula of a compound of Y and chlorine?
g. Is the compound in part f ionic or molecular?

6.66 Consider the following Lewis symbols for elements X and Y:

\[ \text{X} \quad \text{Y} \]

a. What are the group numbers of X and Y?
b. Will a compound of X and Y be ionic or molecular?
c. What ions would be formed by X and Y?
d. What would be the formula of a compound of X and Y?
e. What would be the formula of a compound of X and sulfur?
f. What would be the formula of a compound of Y and chlorine?
g. Is the compound in part f ionic or molecular?

6.67 Using each of the following electron configurations, give the formulas of the cation and anion that form, the formula for the compound they form, and its name. (6.2, 6.3)

<table>
<thead>
<tr>
<th>Electron Configurations</th>
<th>Cation</th>
<th>Anion</th>
<th>Formula of Compound</th>
<th>Name of Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>1s^22s^22p^63s^23p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
</tr>
<tr>
<td>1s^22s^22p^63s^23p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
</tr>
</tbody>
</table>

6.68 Using each of the following electron configurations, give the formulas of the cation and anion that form, the formula for the compound they form, and its name. (6.2, 6.3)

<table>
<thead>
<tr>
<th>Electron Configurations</th>
<th>Cation</th>
<th>Anion</th>
<th>Formula of Compound</th>
<th>Name of Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>1s^22s^22p^63s^23p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
</tr>
<tr>
<td>1s^22s^22p^63s^23p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
<td>1s^22s^22p^6</td>
</tr>
</tbody>
</table>

### ADDITIONAL QUESTIONS AND PROBLEMS

6.69 Write the name for the following: (6.1)

a. N^+ b. Mg^{2+} c. O^2- d. Al^{3+}

6.70 Write the name for the following: (6.1)

a. K^+ b. Na^+ c. Ba^{2+} d. Cl^-

6.71 Consider an ion with the symbol X^{3+} and the electronic configuration 1s^22s^22p^6. (6.1, 6.2, 6.3)

a. What is the group number of the element X?
b. What is the element X?
c. What is the Lewis symbol of this element?
d. What is the formula of the compound formed from X and phosphate?

6.72 Consider an ion with the symbol Z^- and the electronic configuration 1s^22s^22p^63s^23p^6. (6.1, 6.2, 6.3)

a. What is the group number of the element Z?
b. What is the element Z?
c. What is the Lewis symbol of this element?
d. What is the formula of the compound formed from a nickel(II) ion and Z?

6.73 Rust consists of iron(III) oxide and some iron(III) hydroxide. (6.1, 6.2, 6.3, 6.4)

a. What is the symbol of iron(III) ion?
b. How many protons and electrons are there in this ion?
c. What is the formula of iron(III) oxide?
d. What is the formula of iron(III) hydroxide?

6.74 Some ionic compounds such as strontium carbonate and barium chlorate are used as colorant in fireworks. (6.1, 6.2, 6.3, 6.4)

a. What are the symbols of strontium and barium ions?
b. How many protons and electrons are there in a strontium ion?
c. What is the formula of strontium carbonate?
d. What is the formula of barium chlorate?

6.75 Write the formula for each of the following ionic compounds: (6.2, 6.3)

a. silver bromide b. calcium fluoride c. aluminum sulfide d. calcium phosphate e. iron (II) chloride f. magnesium nitride

6.76 Write the formula for each of the following ionic compounds: (6.2, 6.3)

a. nickel(III) oxide b. iron(III) sulfide c. lead(II) sulfate d. chromium(III) iodide e. lithium nitride f. gold(I) oxide

6.77 Name each of the following molecular compounds: (6.5)

a. SF_5 b. PH_3 c. BBF_3
d. PF_5 e. Cl_2O_7 f. P_2O_5

6.78 Name each of the following molecular compounds: (6.5)

a. B_3H_6 b. ClF_3 c. NO_2
d. CCl_4 e. PCl_3 f. N_2O_5
6.79 Write the formula for each of the following molecular compounds: (6.5)
a. carbon monoxide
b. phosphorous pentabromide
c. iodine heptafluoride
d. sulfur trioxide
6.80 Write the formula for each of the following molecular compounds: (6.5)
a. sulfur dibromide b. carbon dioxide
c. nitrogen trichloride d. sulfur tetrafluoride
6.81 Classify each of the following as ionic or molecular, and give its name: (6.3, 6.5)
a. Na₂CO₃ b. NH₃ c. AlBr₃
d. CS₂ e. BN f. Ca₃P₂
6.82 Classify each of the following as ionic or molecular, and give its name: (6.3, 6.5)
a. Al₂(CO₃)₃ b. ClF₅ c. ClO₂
d. Mg₂N₂ e. Ca₃(PO₄)₂
6.83 Write the formula for each of the following: (6.3, 6.4, 6.5)
a. tin(II) carbonate b. lithium phosphide
c. silicon tetrachloride d. manganese(III) oxide
e. tetraphosphorus triselenide
6.84 Write the formula for each of the following: (6.3, 6.4, 6.5)
a. dinitrogen monoxide b. lithium carbonate
c. magnesium hydroxide d. sodium acetate

CHALLENGE QUESTIONS

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

6.85 Complete the following table for atoms or ions: (6.1)

<table>
<thead>
<tr>
<th>Atom or Ion</th>
<th>Number of Protons</th>
<th>Number of Electrons</th>
<th>Electrons Lost/Gained</th>
</tr>
</thead>
<tbody>
<tr>
<td>K⁺</td>
<td>19</td>
<td>18</td>
<td></td>
</tr>
<tr>
<td></td>
<td>12 p⁺ 10 e⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>8 p⁺ 2 e⁻ gained</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>10 e⁻ 3 e⁻ lost</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

6.86 Complete the following table for atoms or ions: (6.1)

<table>
<thead>
<tr>
<th>Atom or Ion</th>
<th>Number of Protons</th>
<th>Number of Electrons</th>
<th>Electrons Lost/Gained</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>30 p⁺ 2 e⁻ lost</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>36 p⁺ 36 e⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>16 p⁺ 2 e⁻ gained</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>46 e⁻ 4 e⁻ lost</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

ANSWERS

Answers to Selected Questions and Problems

6.1 a. 1 b. 2 c. 3 d. 1 e. 2
6.3 a. 2 e⁻ lost b. 3 e⁻ gained c. 1 e⁻ gained
d. 1 e⁻ lost e. 1 e⁻ gained
6.5 a. Li⁺ b. F⁻ c. Mg²⁺ d. Fe³⁺
6.7 a. 29 protons, 27 electrons b. 34 protons, 36 electrons
c. 35 protons, 36 electrons d. 26 protons, 23 electrons
6.9 a. Cl⁻ b. Cs⁺ c. N³⁻ d. Ra²⁺
6.11 a. lithium b. calcium
c. gallium d. phosphide
6.13 a. 8 protons, 10 electrons b. 19 protons, 18 electrons
c. 53 protons, 54 electrons d. 11 protons, 10 electrons
6.15 a and e
d. MgF₂ e. Al₂S₃
6.19  a. K⁺ and S²⁻, K₂S  
   b. Na⁺ and N³⁻, Na₃N  
   c. Al⁺⁺ and I⁻, AlI₃  
   d. Ca²⁺ and O²⁻, Ca₃O₃  
   e. magnesium phosphate  
   f. barium fluoride  

6.21  a. aluminum oxide  
   b. copper(II) oxide  
   c. chromium(III) oxide  
   d. lead(IV) oxide  
   e. manganese(II) oxide  
   f. iron(II) oxide  

6.24  a. tin(II) chloride  
   b. dichlorine heptoxide  
   c. phosphorous pentoxide  
   d. selenium hexafluoride  
   e. selenium hexafluoride  
   f. selenium hexafluoride  

6.25  a. tin(II) chloride  
   b. dichlorine heptoxide  
   c. phosphorous pentoxide  
   d. selenium hexafluoride  
   e. selenium hexafluoride  
   f. selenium hexafluoride  

6.27  a. Au⁺⁺  
   b. Fe⁺⁺  
   c. copper(I) sulfide  
   d. cadmium bromide  
   e. copper(II) sulfate  
   f. copper(II) nitrate  

6.29  a. MgCl₂  
   b. Na₂S  
   c. AuN  
   d. CoF₃  
   e. Cu₃P  
   f. CrCl₂  

6.31  a. CoCl₃  
   b. PbO₂  
   c. AgI  
   d. Ca₃N₂  
   e. FeBr₃  
   f. MgO  

6.33  a. K₂P  
   b. CuCl₂  
   c. PO₃⁻⁻  
   d. ClO₃⁻  
   e. carbonate  
   f. nitrate  

6.37  a. sulfate  
   b. sulfite  
   c. hydrogen sulfite (bisulfite)  

6.39  | NO₂⁻  | CO₃²⁻  | HSO₄⁻  | PO₄²⁻  |
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Li⁺</td>
<td>LiNO₂</td>
<td>Li₂CO₃</td>
<td>LiHSO₄</td>
</tr>
<tr>
<td>Li⁺²</td>
<td>Lithium nitrite</td>
<td>Lithium carbonate</td>
<td>Lithium hydrogen sulfate</td>
</tr>
<tr>
<td>Cu⁺²</td>
<td>Cu(NO₃)₂</td>
<td>CuCO₃</td>
<td>Cu(HSO₄)₂</td>
</tr>
<tr>
<td>Ba⁺²</td>
<td>Ba(NO₃)₂</td>
<td>BaCO₃</td>
<td>Ba(HSO₄)₂</td>
</tr>
</tbody>
</table>

6.61  a. Calcium forms ions by losing electrons to achieve the same electronic configuration as the nearest noble gas. As calcium is in group 2A (2) it loses 2 electrons to become Ca²⁺ ion.  
   b. 1s² 2s² 2p⁶ 3s² 3p⁶  
   c. Ar has the electronic configuration Ca²⁺ (1s² 2s² 2p⁶ 3s² 3p⁶).  

6.63  a. P³⁺ ion  
   b. O atom  
   c. Zn²⁺ ion  
   d. Fe³⁺ ion  

6.65  a. X = Group 1A (1), Y = Group 6A (16)  
   b. H₂O  
   c. X⁺ and Y²⁻  
   d. X₂Y  
   e. X₂S  
   f. YCl₂  
   g. molecular  

6.69  a. nitrogen  
   b. magnesium  
   c. oxide  
   d. aluminum  
   e. dichlorine heptoxide  
   f. phosphorous pentoxide  

6.71  a. 3A (13)  
   b. Al  
   c. Al⁺⁺  
   d. AlPO₄  

6.73  a. Fe³⁺  
   b. 26 protons and 23 electrons  
   c. Fe₂O₃  
   d. Fe(OH)₃  

6.75  a. AgBr  
   b. CaF₂  
   c. Al₂S₃  
   d. Ca₃(PO₄)₂  
   e. FeCl₂  
   f. Mg₃N₂  

6.77  a. sulfur tetrafluoride  
   b. phosphorous trihydride  
   c. boron tribromide  
   d. phosphorous pentfluoride  
   e. dichlorine heptoxide  
   f. phosphorous pentoxide  

6.81  a. ionic, sodium carbonate  
   b. molecular, ammonia  
   c. ionic, aluminum bromide  
   d. molecular, carbon disulfide  
   e. ionic, calcium phosphate  

6.83  a. SnCO₃  
   b. Li₂P  
   c. SiCl₄  
   d. Mn₂O₃  
   e. P₂Se₃  
   f. CaBr₂
### 6.85

<table>
<thead>
<tr>
<th>Atom or Ion</th>
<th>Number of Protons</th>
<th>Number of Electrons</th>
<th>Electrons Lost/Gained</th>
</tr>
</thead>
<tbody>
<tr>
<td>K⁺</td>
<td>19 p⁺</td>
<td>18 e⁻</td>
<td>1 e⁻ lost</td>
</tr>
<tr>
<td>Mg²⁺</td>
<td>12 p⁺</td>
<td>10 e⁻</td>
<td>2 e⁻ lost</td>
</tr>
<tr>
<td>O²⁻</td>
<td>8 p⁺</td>
<td>10 e⁻</td>
<td>2 e⁻ gained</td>
</tr>
<tr>
<td>Al³⁺</td>
<td>13 p⁺</td>
<td>10 e⁻</td>
<td>3 e⁻ lost</td>
</tr>
</tbody>
</table>

### 6.87

a. Li  
b. O  
c. Be

### 6.89

a. ionic, lithium hydrogen phosphate  
b. molecular, chlorine trifluoride  
c. ionic, magnesium chlorite  
d. molecular, nitrogen trifluoride  
e. ionic, calcium bisulfate or calcium hydrogen sulfate  
f. ionic, potassium perchlorate  
g. ionic, gold(III) sulfite
MAX, A SIX-YEAR-OLD dog, is listless, drinking large amounts of water, and not eating his food. His owner takes Max to his veterinarian, Chris, for an examination. Chris weighs Max and obtains a blood sample for a blood chemistry profile, which determines the overall health, detects any metabolic disorders, and measures the concentration of electrolytes.

The results of the lab tests for Max indicate that the white blood count is elevated, which may indicate an infection or inflammation. His electrolytes, which are also indicators of good health, were all in the normal ranges. The electrolyte chloride (Cl⁻) was in the normal range of 0.106 to 0.118 mol/L, and the electrolyte sodium (Na⁺) has a normal range of 0.144 to 0.160 mol/L.

Potassium (K⁺), another important electrolyte, has a normal range of 0.0035 to 0.0058 mol/L. To treat the possible infection, Chris ordered 375-mg tablets of Clavamox, which is a broad-spectrum antibiotic approved for veterinary use in dogs.

CAREER
Veterinarian

Veterinarians care for domesticated pets such as rats, birds, dogs, and cats. Some veterinarians specialize in the treatment of large animals, such as horses and cattle. Veterinarians interact with pet owners as they give advice about feeding, behavior, and breeding. In the assessment of an ill animal, they record the animal’s symptoms and medical history including dietary intake, medications, eating habits, weight, and any signs of disease.

To diagnose health problems, veterinarians perform laboratory tests on animals including a complete blood count and urinalysis. They also obtain tissue and blood samples, as well as vaccinate against distemper, rabies, and other diseases. Veterinarians prescribe medication if an animal has an infection or illness. If an animal is injured, a veterinarian treats wounds, and sets fractures. A veterinarian also performs surgery such as neutering and spaying, provides dental cleanings, removes tumors, and euthanizes animals.
**7.1 The Mole**

**LEARNING GOAL** Use Avogadro’s number to calculate the number of particles in a given number of moles. Calculate the number of moles of an element in a given number of moles of a compound.

At the grocery store, you buy eggs by the dozen or soda by the case. In an office-supply store, pencils are ordered by the gross and paper by the ream. Common terms such as **dozen, case, gross, and ream** are used to count the number of items present. For example, when you buy a dozen eggs, you know you will get 12 eggs in the carton.

**Avogadro’s Number**

In chemistry, particles such as atoms, molecules, and ions are counted by the **mole** (abbreviated **mol** in calculations), which contains $6.022 \times 10^{23}$ items. This value, known as **Avogadro’s number**, is a very big number because atoms are so small that it takes an extremely large number of atoms to provide a sufficient amount to weigh and use in chemical reactions. Avogadro’s number is named for Amedeo Avogadro (1776–1856), an Italian physicist.

Collections of items include dozen, case, gross, and ream.

![Image of items: 3 cases of soda, 24 cans per case.](image)

24 cans = 1 case

144 pencils = 1 gross

500 sheets = 1 ream

12 eggs = 1 dozen

$6.022 \times 10^{23} = 602\,200\,000\,000\,000\,000\,000\,000$

One mole of any element always contains Avogadro’s number of atoms. For example, 1 mol of carbon contains $6.022 \times 10^{23}$ carbon atoms; 1 mol of aluminum contains $6.022 \times 10^{23}$ aluminum atoms; 1 mol of sulfur contains $6.022 \times 10^{23}$ sulfur atoms.

$1\,\text{mol of an element} = 6.022 \times 10^{23}\,\text{atoms of that element}$

Avogadro’s number tells us that 1 mol of a compound contains $6.022 \times 10^{23}$ of the particular type of particles that make up that compound. One mole of a molecular...
compound contains Avogadro’s number of molecules. For example, 1 mol of CO₂ contains $6.022 \times 10^{23}$ molecules of CO₂. For an ionic compound, 1 mol contains Avogadro’s number of formula units, which are the groups of ions represented by its formula. For the ionic formula, NaCl, 1 mol contains $6.022 \times 10^{23}$ formula units of NaCl (Na⁺, Cl⁻).

**TABLE 7.1** gives examples of the number of particles in some 1-mol quantities.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Number and Type of Particles</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mol of Al</td>
<td>$6.022 \times 10^{23}$ atoms of Al</td>
</tr>
<tr>
<td>1 mol of Fe</td>
<td>$6.022 \times 10^{23}$ atoms of Fe</td>
</tr>
<tr>
<td>1 mol of water (H₂O)</td>
<td>$6.022 \times 10^{23}$ molecules of H₂O</td>
</tr>
<tr>
<td>1 mol of vitamin C (C₆H₈O₆)</td>
<td>$6.022 \times 10^{24}$ molecules of vitamin C</td>
</tr>
<tr>
<td>1 mol of NaCl</td>
<td>$6.022 \times 10^{23}$ formula units of NaCl</td>
</tr>
</tbody>
</table>

**Using Avogadro’s Number as a Conversion Factor**

We can use Avogadro’s number as a conversion factor to convert between the moles of a substance and the number of particles it contains.

$$\frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mol}} \quad \text{and} \quad \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ particles}}$$

For example, we use Avogadro’s number to convert 4.00 mol of iron to atoms of iron.

$$4.00 \text{ mol Fe} \times \frac{6.022 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} = 2.41 \times 10^{24} \text{ atoms of Fe}$$

**ENGAGE**

Why is 0.20 mol of aluminum a small number, but the number of atoms in 0.20 mol is a large number: $1.2 \times 10^{23}$ atoms of aluminum?

We can also use Avogadro’s number to convert $3.01 \times 10^{24}$ molecules of CO₂ to moles of CO₂.

$$3.01 \times 10^{24} \text{ molecules CO}_2 \times \frac{1 \text{ mol CO}_2}{6.022 \times 10^{23} \text{ molecules CO}_2} = 5.00 \text{ mol of CO}_2$$

In calculations that convert between moles and particles, the number of moles will be small compared to the number of atoms or molecules, which will be large as shown in Sample Problem 7.1.

**SAMPLE PROBLEM 7.1 Calculating the Number of Molecules**

How many molecules are present in 1.75 mol of carbon dioxide?

The solid form of carbon dioxide is known as “dry ice.”
7.1 The Mole

The Mole

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.75 mol of CO₂</td>
<td>molecules of CO₂</td>
<td>Avogadro’s number</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to convert moles to particles.

moles of CO₂ Avogadro’s number molecules of CO₂

STEP 3 Use Avogadro’s number to write conversion factors.

\[
\frac{6.022 \times 10^{23} \text{ molecules CO}_2}{1 \text{ mol CO}_2} \quad \text{and} \quad \frac{1 \text{ mol CO}_2}{6.022 \times 10^{23} \text{ molecules CO}_2}
\]

STEP 4 Set up the problem to calculate the number of particles.

\[
1.75 \text{ mol CO}_2 \times \frac{6.022 \times 10^{23} \text{ molecules CO}_2}{1 \text{ mol CO}_2} = 1.05 \times 10^{24} \text{ molecules of CO}_2
\]

STUDY CHECK 7.1

How many moles of water, H₂O, contain \(2.60 \times 10^{23}\) molecules of water?

ANSWER

0.432 mol of H₂O

Moles of Elements in a Chemical Compound

We have seen that the subscripts in a chemical formula indicate the number of atoms of each type of element in a compound. For example, aspirin, C₉H₈O₄, is a drug used to reduce pain and inflammation in the body. Using the subscripts in the chemical formula of aspirin shows that there are 9 carbon atoms, 8 hydrogen atoms, and 4 oxygen atoms. The subscripts of the formula of aspirin, C₉H₈O₄, also tell us the number of moles of each element in 1 mol of aspirin: 9 mol of C atoms, 8 mol of H atoms, and 4 mol of O atoms.

Why does 1 mol of Zn(C₂H₅O₂)₂, a dietary supplement, contain 1 mol of Zn, 4 mol of C, 6 mol of H, and 4 mol of O?

ENGAGE

The chemical formula subscripts specify the

- Atoms in 1 molecule
  - C₉H₈O₄
  - Carbon: 9 atoms of C | Hydrogen: 8 atoms of H | Oxygen: 4 atoms of O
- Moles of each element in 1 mol
  - 9 mol of C | 8 mol of H | 4 mol of O
Using a Chemical Formula to Derive Conversion Factors

Using the subscripts from the formula, C₉H₈O₄, we can write the conversion factors for each of the elements in 1 mol of aspirin:

<table>
<thead>
<tr>
<th>Element</th>
<th>Conversion Factor</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>9 mol C / 1 mol C₉H₈O₄</td>
</tr>
<tr>
<td>H</td>
<td>8 mol H / 1 mol C₉H₈O₄</td>
</tr>
<tr>
<td>O</td>
<td>4 mol O / 1 mol C₉H₈O₄</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 7.2 Calculating the Moles of an Element in a Compound**

How many moles of carbon are present in 1.50 mol of aspirin, C₉H₈O₄?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.50 mol of aspirin, C₉H₈O₄</td>
<td></td>
<td>moles of C</td>
<td>subscripts in formula</td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to convert moles of a compound to moles of an element.

moles of C₉H₈O₄ \( \text{Subscript} \) moles of C

**STEP 3** Write the equalities and conversion factors using subscripts.

\[
\frac{9 \text{ mol C}}{1 \text{ mol C₉H₈O₄}} = \frac{9 \text{ mol C}}{1 \text{ mol C₉H₈O₄}} \quad \text{and} \quad \frac{1 \text{ mol C₉H₈O₄}}{9 \text{ mol C}}
\]

**STEP 4** Set up the problem to calculate the moles of an element.

\[
1.50 \text{ mol C₉H₈O₄} \times \frac{9 \text{ mol C}}{1 \text{ mol C₉H₈O₄}} = 13.5 \text{ mol C}
\]

**STUDY CHECK 7.2**

How many moles of aspirin, C₉H₈O₄, contain 0.480 mol of O?

**ANSWER**

0.120 mol of aspirin

**QUESTIONS AND PROBLEMS**

**7.1 The Mole**

**LEARNING GOAL** Use Avogadro’s number to calculate the number of particles in a given number of moles. Calculate the number of moles of an element in a given number of moles of a compound.

**7.1** What is a mole?

**7.2** What is Avogadro’s number?

**7.3** Calculate each of the following:

- a. number of C atoms in 0.500 mol of C
- b. number of SO₂ molecules in 1.28 mol of SO₂
- c. moles of Fe in \( 5.22 \times 10^{22} \) atoms of Fe
- d. moles of C₂H₆O in \( 8.50 \times 10^{24} \) molecules of C₂H₆O

**7.4** Calculate each of the following:

- a. number of Li atoms in 4.5 mol of Li
- b. number of CO₂ molecules in 0.0180 mol of CO₂
- c. moles of Cu in \( 7.8 \times 10^{21} \) atoms of Cu
- d. moles of C₂H₆ in \( 3.75 \times 10^{23} \) molecules of C₂H₆

**7.5** Calculate each of the following quantities in 2.00 mol of H₃PO₄:

- a. moles of H
- b. moles of O
- c. atoms of P
- d. atoms of O

**7.6** Calculate each of the following quantities in 0.185 mol of C₆H₁₄O:

- a. moles of C
- b. moles of O
- c. atoms of H
- d. atoms of C
Applications

7.7 Quinine, C_{20}H_{24}N_{2}O_{2}, is a component of tonic water and bitter lemon.
   a. How many moles of H are in 1.5 mol of quinine?
   b. How many moles of C are in 5.0 mol of quinine?
   c. How many moles of N are in 0.020 mol of quinine?

7.8 Aluminum sulfate, Al_{2}(SO_{4})_{3}, is used in some antiperspirants.
   a. How many moles of O are present in 3.0 mol of Al_{2}(SO_{4})_{3}?
   b. How many moles of aluminum ions (Al^{3+}) are present in 0.40 mol of Al_{2}(SO_{4})_{3}?
   c. How many moles of sulfate ions (SO_{4}^{2-}) are present in 1.5 mol of Al_{2}(SO_{4})_{3}?

7.9 Naproxen is used to treat pain and inflammation caused by arthritis. Naproxen has a formula of C_{14}H_{14}O_{3}.
   a. How many moles of C are present in 2.30 mol of naproxen?
   b. How many moles of H are present in 0.444 mol of naproxen?
   c. How many moles of O are present in 0.0765 mol of naproxen?

7.10 Benadryl is an over-the-counter drug used to treat allergy symptoms. The formula of Benadryl is C_{17}H_{21}NO.
   a. How many moles of C are present in 0.733 mol of Benadryl?
   b. How many moles of H are present in 2.20 mol of Benadryl?
   c. How many moles of N are present in 1.54 mol of Benadryl?

7.2 Molar Mass

LEARNING GOAL Given the chemical formula of a substance, calculate its molar mass.

A single atom or molecule is much too small to weigh, even on the most accurate balance. In fact, it takes a huge number of atoms or molecules to make enough of a substance for you to see. An amount of water that contains Avogadro’s number of water molecules is only a few sips. However, in the laboratory, we can use a balance to weigh out Avogadro’s number of particles for 1 mol of substance.

For any element, the quantity called molar mass is the quantity in grams that equals the atomic mass of that element. We are counting \(6.022 \times 10^{23}\) atoms of an element when we weigh out the number of grams equal to its molar mass. For example, carbon has an atomic mass of 12.01 on the periodic table. This means 1 mol of carbon atoms has a mass of 12.01 g. Then to obtain 1 mol of carbon atoms, we would need to weigh out 12.01 g of carbon. Thus, the molar mass of carbon is found by looking at its atomic mass on the periodic table.
Molar Mass of a Compound

To determine the molar mass of a compound, multiply the molar mass of each element by its subscript in the formula and add the results as shown in Sample Problem 7.3. In this text, we round the molar mass of an element to the hundredths place (0.01) or use at least four significant figures for calculations.

**SAMPLE PROBLEM 7.3 Calculating the Molar Mass of a Compound**

Calculate the molar mass for lithium carbonate, \( \text{Li}_2\text{CO}_3 \), used to treat bipolar disorder.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>formula Li₂CO₃</td>
<td>molar mass of Li₂CO₃</td>
<td>periodic table</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Obtain the molar mass of each element.

<table>
<thead>
<tr>
<th>Grams from 2 mol of Li</th>
<th>Grams from 1 mol of C</th>
<th>Grams from 3 mol of O</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \frac{2 \text{ mol Li}}{1 \text{ mol Li}} \times \frac{6.941 \text{ g Li}}{1 \text{ mol Li}} ) = 13.88 g of Li</td>
<td>( \frac{1 \text{ mol C}}{1 \text{ mol C}} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} ) = 12.01 g of C</td>
<td>( \frac{3 \text{ mol O}}{1 \text{ mol O}} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} ) = 48.00 g of O</td>
</tr>
</tbody>
</table>

**STEP 2** Multiply each molar mass by the number of moles (subscript) in the formula.

**STEP 3** Calculate the molar mass by adding the masses of the elements.

\[
\begin{align*}
\text{Molar mass of Li}_2\text{CO}_3 &= 13.88 \text{ g of Li} + 12.01 \text{ g of C} + 48.00 \text{ g of O} \\
&= 73.89 \text{ g}
\end{align*}
\]

**STUDY CHECK 7.3**

Calculate the molar mass for salicylic acid, \( \text{C}_7\text{H}_6\text{O}_3 \), which is used to treat skin conditions such as acne, psoriasis, and dandruff.

**ANSWER**

138.12 g
FIGURE 7.1 shows some 1-mol quantities of substances.

1-Mol Quantities

- S
- Fe
- NaCl
- K₂Cr₂O₇
- C₁₂H₂₂O₁₁

FIGURE 7.1 1-mol samples: sulfur, S (32.07 g); iron, Fe (55.85 g); salt, NaCl (58.44 g); potassium dichromate, K₂Cr₂O₇ (294.2 g); and sucrose, C₁₂H₂₂O₁₁ (342.3 g).

How is the molar mass for K₂Cr₂O₇ obtained?

QUESTIONS AND PROBLEMS

7.2 Molar Mass

LEARNING GOAL Given the chemical formula of a substance, calculate its molar mass.

7.11 Calculate the molar mass for each of the following:
   a. Cl₂
   b. C₂H₆O₃
   c. Mg₃(PO₄)₂

7.12 Calculate the molar mass for each of the following:
   a. O₂
   b. K₂HPO₄
   c. Fe(CIO₄)₃

7.13 Calculate the molar mass for each of the following:
   a. AlF₃
   b. C₂H₆Cl₂
   c. SnF₂

7.14 Calculate the molar mass for each of the following:
   a. C₂H₅O₄
   b. Ga₃(CO₃)₃
   c. KBrO₄

Applications

7.15 Calculate the molar mass for each of the following:
   a. NaCl, table salt
   b. Fe₂O₃, rust
   c. C₁₀H₂₀FNO₃, Paxil, an antidepressant

7.16 Calculate the molar mass for each of the following:
   a. FeSO₄, iron supplement
   b. Al₂O₃, absorbent and abrasive
   c. C₅H₄NO₃, saccharin, an artificial sweetener

7.17 Calculate the molar mass for each of the following:
   a. Al₂(SO₄)₃, antiperspirant
   b. KC₄H₄O₆, cream of tartar
   c. C₁₈H₁₉N₄O₇S, amoxicillin, an antibiotic

7.18 Calculate the molar mass for each of the following:
   a. C₃H₇O₂, rubbing alcohol
   b. (NH₄)₂CO₃, baking powder
   c. Zn(C₂H₃O₂)₂, a zinc supplement

7.19 Calculate the molar mass for each of the following:
   a. C₅H₄NO₂, acetaminophen used in Tylenol
   b. C₉(C₅H₄O₂)₂, a calcium supplement
   c. C₁₇H₁₈FN₃O₃, Cipro, used to treat a range of bacterial infections

7.20 Calculate the molar mass for each of the following:
   a. CaSO₄, calcium sulfate, used to make casts to protect broken bones
   b. C₉H₁₂N₂O₂Pt, Carboplatin, used in chemotherapy
   c. C₁₂H₁₈O₃, Propofol, used to induce anesthesia during surgery

7.3 Calculations Using Molar Mass

LEARNING GOAL Given the number of moles (or grams) of a substance, calculate the grams (or moles). Given the grams of a compound, calculate the grams of one of the elements.

The molar mass of an element is one of the most useful conversion factors in chemistry because it converts moles of a substance to grams, or grams to moles. For example, 1 mol of silver has a mass of 107.9 g. To express molar mass of Ag as an equality, we write

$$1 \text{ mol of Ag} = 107.9 \text{ g of Ag}$$
From this equality for a molar mass, two conversion factors can be written as

\[
\begin{align*}
107.9 \text{ g Ag} & \quad \text{and} \quad 1 \text{ mol Ag} \\
1 \text{ mol Ag} & \quad \text{and} \quad 107.9 \text{ g Ag}
\end{align*}
\]

Sample Problem 7.4 shows how the molar mass of silver is used as a conversion factor.

**SAMPLE PROBLEM 7.4 Converting Moles of an Element to Grams**

Silver metal is used in the manufacture of tableware, mirrors, jewelry, and dental alloys. If the design for a piece of jewelry requires 0.750 mol of silver, how many grams of silver are needed?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.750 mol of Ag</td>
<td>grams of Ag</td>
<td>molar mass</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to convert moles to grams.

\[
\text{moles of Ag} \quad \text{Molar mass} \quad \text{grams of Ag}
\]

**STEP 3** Determine the molar mass and write conversion factors.

\[
\begin{align*}
1 \text{ mol Ag} & = \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} \\
1 \text{ mol Ag} & = \frac{1 \text{ mol Ag}}{107.9 \text{ g Ag}}
\end{align*}
\]

**STEP 4** Set up the problem to convert moles to grams.

\[
0.750 \text{ mol Ag} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 80.9 \text{ g of Ag}
\]

**STUDY CHECK 7.4**

A dentist orders 24.4 g of gold (Au) to prepare dental crowns and fillings. Calculate the number of moles of gold in the order.

**ANSWER**

0.124 mol of Au

**Writing Conversion Factors for the Molar Mass of a Compound**

The conversion factors for a compound are also written from the molar mass. For example, the molar mass of the compound H$_2$O is written

\[
1 \text{ mol of H}_2\text{O} = 18.02 \text{ g of H}_2\text{O}
\]

From this equality, conversion factors for the molar mass of H$_2$O are written as

\[
\begin{align*}
18.02 \text{ g H}_2\text{O} & \quad \text{and} \quad 1 \text{ mol H}_2\text{O} \\
1 \text{ mol H}_2\text{O} & \quad \text{and} \quad 18.02 \text{ g H}_2\text{O}
\end{align*}
\]

We can now change from moles to grams, or grams to moles, using the conversion factors derived from the molar mass of a compound shown in Sample Problem 7.5. (Remember, you must determine the molar mass first.)
SAMPLE PROBLEM 7.5 Converting Mass of a Compound to Moles
A box of salt contains 737 g of NaCl. How many moles of NaCl are present in the box?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>737 g NaCl</td>
<td>moles NaCl</td>
<td>molar mass</td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to convert grams to moles.

grams of NaCl $\rightarrow$ Molar mass $\rightarrow$ moles of NaCl

STEP 3 Determine the molar mass and write conversion factors.

\[
(1 \times 22.99) + (1 \times 35.45) = 58.44 \text{ g/mol}
\]

\[
\begin{align*}
1 \text{ mol of NaCl} &= 58.44 \text{ g of NaCl} \\
58.44 \text{ g NaCl} &= 1 \text{ mol NaCl} \\
&= 58.44 \text{ g NaCl}
\end{align*}
\]

STEP 4 Set up the problem to convert grams to moles.

\[
737 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 12.6 \text{ mol NaCl}
\]

STUDY CHECK 7.5

One tablet of an antacid contains 680 mg of CaCO₃. How many moles of CaCO₃ are present?

ANSWER

0.00679 or $6.79 \times 10^{-3}$ mol of CaCO₃

FIGURE 7.2 gives a summary of the calculations to show the connections between the moles of a compound, its mass in grams, the number of molecules (or formula units if ionic), and the moles and atoms of each element in that compound in the following flowchart:

Why are there more grams of chlorine than grams of fluorine in 1 mol of Freon-12, $\text{CCl}_2\text{F}_2$?

FIGURE 7.2 The moles of a compound are related to its mass in grams by molar mass, to the number of molecules (or formula units) by Avogadro’s number, and to the moles of each element by the subscripts in the formula.

What steps are needed to calculate the number of atoms of H in 5.00 g of $\text{CH}_4$?
We can now convert the mass in grams of a compound to the mass of one of the elements in a compound as shown in Sample Problem 7.6.

**SAMPLE PROBLEM 7.6 Converting Grams of Compound to Grams of Element**

Hot packs are used to reduce muscle aches, inflammation, and muscle spasms. A hot pack consists of a bag of water and an inner bag containing 10.2 g of CaCl₂. When the bag is smashed, the CaCl₂ dissolves in the water and heat is released. How many grams of Cl are in the CaCl₂ in the inner bag?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>10.2 g of CaCl₂</td>
<td>grams of Cl</td>
<td>molar mass, subscript for Cl</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to convert grams of a compound to grams of an element.

grams of CaCl₂  Molar mass  moles of CaCl₂  Mole factor  moles of Cl  Molar mass  grams of Cl

**STEP 3** Write the equalities and conversion factors for molar mass and mole factors.

\[
\begin{align*}
1 \text{ mol of CaCl}_2 & = \frac{110.98 \text{ g of CaCl}_2}{1 \text{ mol CaCl}_2} \\
2 \text{ mol Cl} & = \frac{1 \text{ mol CaCl}_2}{1 \text{ mol CaCl}_2} \\
1 \text{ mol of Cl} & = \frac{35.45 \text{ g of Cl}}{1 \text{ mol Cl}}
\end{align*}
\]

**STEP 4** Set up the problem to convert grams of a compound (or element) to grams of an element (or compound).

\[
10.2 \text{ g CaCl}_2 \times \frac{1 \text{ mol CaCl}_2}{110.98 \text{ g CaCl}_2} \times \frac{2 \text{ mol Cl}}{1 \text{ mol CaCl}_2} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 6.52 \text{ g of Cl}
\]

**STUDY CHECK 7.6**

Tin(II) fluoride (SnF₂) is added to toothpaste to strengthen tooth enamel. How many grams of tin(II) fluoride contain 1.46 g of F?

**ANSWER**

6.02 g of SnF₂
7.4 Mass Percent Composition

**LEARNING GOAL** Given the formula of a compound, calculate the mass percent composition.

Because the atoms of the elements in a compound are combined in a definite mole ratio, they are also combined in a definite proportion by mass. When we know the mass of an element in the mass of a sample of a compound, we can calculate its **mass percent composition** or **mass percent**, which is the mass of an element divided by the total mass of the compound and multiplied by 100%. For example, we can calculate the mass percent of N.
if we find from experiment that 7.64 g of N are present in 12.0 g of N₂O, “laughing gas,” which is used as an anesthetic for surgery and in dentistry.

From the grams of N and the grams of N₂O, we calculate the mass percent of nitrogen as follows:

\[
\text{Mass percent of an element} = \frac{\text{mass of an element}}{\text{total mass of the compound}} \times 100%
\]

\[
\text{Mass percent of N} = \frac{7.64 \text{ g N}}{12.0 \text{ g N}_2\text{O}} \times 100% = 63.7% \text{ N}
\]

**Core Chemistry Skill**
Calculating Mass Percent Composition

**SAMPLE PROBLEM 7.7** Calculating Mass Percent Composition from Molar Mass

The odor of pears is due to the compound propyl acetate, which has a formula of C₅H₁₀O₂. What is its mass percent composition?

**TRY IT FIRST**

**SOLUTION**

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>C₅H₁₀O₂</td>
<td>mass percent composition: %C, %H, %O</td>
<td>periodic table</td>
</tr>
</tbody>
</table>

**STEP 1** Determine the total mass of each element in the molar mass of a formula.

\[
5 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 60.05 \text{ g of C}
\]

\[
10 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 10.08 \text{ g of H}
\]

\[
2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g of O}
\]

Molar mass of C₅H₁₀O₂ = 102.13 g of C₅H₁₀O₂

**STEP 2** Divide the total mass of each element by the molar mass and multiply by 100%.

\[
\text{Mass } % \text{ C} = \frac{60.05 \text{ g C}}{102.13 \text{ g } C_5H_{10}O_2} \times 100% = 58.80% \text{ C}
\]

\[
\text{Mass } % \text{ H} = \frac{10.08 \text{ g H}}{102.13 \text{ g } C_5H_{10}O_2} \times 100% = 9.870% \text{ H}
\]

\[
\text{Mass } % \text{ O} = \frac{32.00 \text{ g O}}{102.13 \text{ g } C_5H_{10}O_2} \times 100% = 31.33% \text{ O}
\]

The total mass percent for all the elements in the compound should equal 100%. In some cases, because of rounding off, the sum of the mass percents may not total exactly 100%.

58.80% C + 9.870% H + 31.33% O = 100.00%

**STUDY CHECK 7.7**
Ethylene glycol, C₂H₆O₂, used as automobile antifreeze, is a sweet-tasting liquid, which is toxic to humans and animals. What is the mass percent composition of ethylene glycol?

**ANSWER**
38.70% C; 9.744% H; 51.56% O
Every year in the spring, homeowners and farmers add fertilizers to the soil to produce greener lawns and larger crops. Plants require several nutrients, but the major ones are nitrogen, phosphorus, and potassium. Nitrogen promotes green growth, phosphorus promotes strong root development for strong plants and abundant flowers, and potassium helps plants defend against diseases and weather extremes. The numbers on a package of fertilizer give the percentages each of N, P, and K by mass. For example, the set of numbers 30–3–4 describes a fertilizer that contains 30% N, 3% P, and 4% K.

The major nutrient, nitrogen, is present in huge quantities as N₂ in the atmosphere, but plants cannot utilize nitrogen in this form. Bacteria in the soil convert atmospheric N₂ to usable forms by nitrogen fixation. To provide additional nitrogen to plants, several types of nitrogen-containing chemicals, including ammonia, nitrates, and ammonium compounds, are added to soil. The nitrates are absorbed directly, but ammonia and ammonium salts are first converted to nitrates by the soil bacteria.

The percent nitrogen depends on the type of nitrogen compound used in the fertilizer. The percent nitrogen by mass in each type is calculated using mass percent composition.

### Type of Fertilizer | Percent Nitrogen by Mass
---|---
NH₃ | \( \frac{14.01 \text{ g N}}{17.03 \text{ g NH}_3} \times 100\% = 82.27\% \text{ N} \)
NH₄NO₃ | \( \frac{28.02 \text{ g N}}{80.05 \text{ g NH}_4\text{NO}_3} \times 100\% = 35.00\% \text{ N} \)
\((\text{NH}_4)_2\text{HPO}_4\) | \( \frac{28.02 \text{ g N}}{132.06 \text{ g (NH}_4)_2\text{HPO}_4} \times 100\% = 21.22\% \text{ N} \)
\((\text{NH}_4)_2\text{SO}_4\) | \( \frac{28.02 \text{ g N}}{132.15 \text{ g (NH}_4)_2\text{SO}_4} \times 100\% = 21.20\% \text{ N} \)

The choice of a fertilizer depends on its use and convenience. A fertilizer can be prepared as crystals or a powder, in a liquid solution, or as a gas such as ammonia. The ammonia and ammonium fertilizers are water soluble and quick-acting. Other forms may be made to slow-release by enclosing water-soluble ammonium salts in a thin plastic coating. The most commonly used fertilizer is NH₄NO₃ because it is easy to apply and has a high percent of N by mass.

### 7.4 Mass Percent Composition

**LEARNING GOAL** Given the formula of a compound, calculate the mass percent composition.

**7.37** Calculate the mass percent composition of each of the following:
- a. 4.68 g of Si and 5.32 g of O
- b. 5.72 g of C and 1.28 g of H
- c. 16.1 g of Na, 22.5 g of S, and 11.3 g of O
- d. 6.22 g of C, 1.04 g of H, and 4.14 g of O

**7.38** Calculate the mass percent composition of each of the following:
- a. 0.890 g of Ba and 1.04 g of Br
- b. 3.82 g of K and 1.18 g of I
- c. 3.29 g of N, 0.946 g of H, and 3.76 g of S
- d. 4.14 g of C, 0.695 g of H, and 2.76 g of O

**Applications**

**7.39** Calculate the mass percent composition of each of the following:
- a. MgF₂, magnesium fluoride
- b. Ca(OH)₂, calcium hydroxide
- c. C₆H₁₂O₆, erythrose, a carbohydrate
- d. (NH₄)₃PO₄, ammonium phosphate, fertilizer
- e. C₁₇H₁₉NO₃, morphine, a painkiller

**7.40** Calculate the mass percent composition of each of the following:
- a. CaCl₂, calcium chloride
- b. Na₂Cr₂O₇, sodium dichromate
- c. C₂H₁₀Cl₃, trichloroethane, a cleaning solvent
- d. C₆H₁₅N₂O₄, calcium phosphate, found in bone and teeth
- e. C₁₈H₉₀O₂, stearic acid, a fatty acid

**7.41** Calculate the mass percent of N in each of the following:
- a. N₂O₅, dinitrogen pentoxide
- b. NH₄Cl, expectorant in cough medicine
- c. C₂H₈N₂, dimethylhydrazine, rocket fuel
- d. C₉H₁₃N₂O₅, Rogaine, stimulates hair growth
- e. C₁₄H₂₂N₂O₇, lidocaine, local anesthetic

**7.42** Calculate the mass percent of S in each of the following:
- a. Na₂SO₄, sodium sulfate
- b. Al₂S₃, aluminum sulfide
- c. SO₃, sulfur trioxide
- d. C₂H₄SO₂, dimethylosulfoxide, topical anti-inflammatory
- e. C₁₀H₁₈N₂O₅S, sulfadiazine, antibacterial
7.5 Empirical Formulas

LEARNING GOAL From the mass percent composition, calculate the empirical formula for a compound.

Up to now, the formulas you have seen have been molecular formulas, which are the actual formulas of compounds. If we write a formula that represents the lowest whole-number ratio of the atoms in a compound, it is called the simplest or empirical formula. For example, the compound benzene, with molecular formula C₆H₆, has the empirical formula CH. Some molecular formulas and their empirical formulas are shown in Table 7.2.

<table>
<thead>
<tr>
<th>Name</th>
<th>Molecular (actual formula)</th>
<th>Empirical (simplest formula)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetylene</td>
<td>C₂H₂</td>
<td>CH</td>
</tr>
<tr>
<td>Benzene</td>
<td>C₆H₆</td>
<td>CH</td>
</tr>
<tr>
<td>Ammonia</td>
<td>NH₃</td>
<td>NH₁</td>
</tr>
<tr>
<td>Hydrazine</td>
<td>N₂H₄</td>
<td>NH₂</td>
</tr>
<tr>
<td>Ribose</td>
<td>C₅H₁₀O₅</td>
<td>CH₂O</td>
</tr>
<tr>
<td>Glucose</td>
<td>C₆H₁₂O₆</td>
<td>CH₂O</td>
</tr>
</tbody>
</table>

The empirical formula of a compound is determined by converting the number of grams of each element to moles and finding the lowest whole-number ratio to use as subscripts, as shown in Sample Problem 7.8.

SAMPLE PROBLEM 7.8 Calculating an Empirical Formula

A compound of iron and chlorine is used to purify water in water-treatment plants. What is the empirical formula of the compound if experimental analysis shows that a sample of the compound contains 6.87 g of iron and 13.1 g of chlorine?

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>6.87 g of Fe</td>
<td>empirical formula</td>
<td>molar mass</td>
</tr>
<tr>
<td></td>
<td>13.1 g of Cl</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

STEP 1 Calculate the moles of each element.

\[
6.87 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 0.123 \text{ mol of Fe}
\]

\[
13.1 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.370 \text{ mol of Cl}
\]

STEP 2 Divide by the smallest number of moles. For this problem, the smallest number of moles is 0.123.

\[
\frac{0.123 \text{ mol Fe}}{0.123} = 1.00 \text{ mol of Fe}
\]

\[
\frac{0.370 \text{ mol Cl}}{0.123} = 3.01 \text{ mol of Cl}
\]
STEP 3 Use the lowest whole-number ratio of moles as subscripts. The relationship of moles of Fe to moles of Cl is 1 to 3, which we obtain by rounding off 3.01 to 3.

\[ \text{Fe}_{1.00} \text{Cl}_{3.01} \rightarrow \text{Fe}_1 \text{Cl}_3, \text{ written as } \text{FeCl}_3 \quad \text{Empirical formula} \]

STUDY CHECK 7.8
Phosphine is a highly toxic compound used for pest and rodent control. If a sample of phosphine contains 0.456 g of P and 0.0440 g of H, what is its empirical formula?

**ANSWER**
\[ \text{PH}_3 \]

Often, the mass percent of each element in a compound is given. The mass percent composition is true for any quantity of the compound. For example, methane, \( \text{CH}_4 \), always has a mass percent composition of 74.9% C and 25.1% H. Thus, if we assume that we have a sample of 100. g of the compound, we can determine the grams of each element in that 100.-g sample and use these to calculate the empirical formula as shown in Sample Problem 7.9.

**SAMPLE PROBLEM 7.9 Calculating an Empirical Formula from the Mass Percent Composition**
Tetrachloroethene is a colorless liquid used for dry cleaning. What is the empirical formula of tetrachloroethene if its mass percent composition is 14.5% C and 85.5% Cl?

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>14.5% C, 85.5% Cl</td>
<td>empirical formula</td>
<td>molar mass</td>
<td></td>
</tr>
</tbody>
</table>
STEP 1 Calculate the moles of each element. In a sample size of 100. g of this compound, there are 14.5 g of C and 85.5 g of Cl.

\[
\begin{align*}
14.5 \text{ g C} & \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.21 \text{ mol of C} \\
85.5 \text{ g Cl} & \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 2.41 \text{ mol of Cl}
\end{align*}
\]

STEP 2 Divide by the smallest number of moles. For this problem, the smallest number of moles is 1.21.

\[
\begin{align*}
\frac{1.21 \text{ mol C}}{1.21} &= 1.00 \text{ mol of C} \\
\frac{2.41 \text{ mol Cl}}{1.21} &= 1.99 \text{ mol of Cl}
\end{align*}
\]

STEP 3 Use the lowest whole-number ratio of moles as subscripts.

\[
\text{C}_{1.00}\text{Cl}_{1.99} \rightarrow \text{C}_1\text{Cl}_2 = \text{CCl}_2
\]

**STUDY CHECK 7.9**

Sulfate of potash is the common name of a compound used in fertilizers to supply potassium and sulfur. What is the empirical formula of this compound if it has a mass percent composition of 44.9% K, 18.4% S, and 36.7% O?

**ANSWER**

\[\text{K}_2\text{SO}_4\]

**Converting Decimal Numbers to Whole Numbers**

Sometimes the result of dividing by the smallest number of moles gives a decimal instead of a whole number. Decimal values that are very close to whole numbers can be rounded off. For example, 2.04 rounds off to 2 and 6.98 rounds off to 7. However, a decimal that is greater than 0.1 or less than 0.9 should not be rounded off. Instead, we multiply by a small integer to obtain a whole number. Some multipliers that are typically used are listed in **TABLE 7.3**.

Let us suppose the numbers of moles we obtain give subscripts in the ratio of \(\text{C}_{1.00}\text{H}_{2.33}\text{O}_{0.99}\). While 0.99 rounds off to 1, we cannot round off 2.33. If we multiply 2.33 \(\times\) 2, we obtain 4.66, which is still not a whole number. If we multiply 2.33 by 3, the answer is 6.99, which rounds off to 7. To complete the empirical formula, **all the other subscripts must be multiplied by 3**.

\[
\text{C}_{1(1.00 \times 3)}\text{H}_{(2.33 \times 3)}\text{O}_{(0.99 \times 3)} = \text{C}_{3.00}\text{H}_{6.99}\text{O}_{2.97} \rightarrow \text{C}_3\text{H}_7\text{O}_3
\]

**TABLE 7.3** Some Multipliers That Convert Decimals to Whole-Number Subscripts

<table>
<thead>
<tr>
<th>Decimal</th>
<th>Multiplier</th>
<th>Example</th>
<th>Whole Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.20</td>
<td>5</td>
<td>1.20 (\times) 5 = 6</td>
<td></td>
</tr>
<tr>
<td>0.25</td>
<td>4</td>
<td>2.25 (\times) 4 = 9</td>
<td></td>
</tr>
<tr>
<td>0.33</td>
<td>3</td>
<td>1.33 (\times) 3 = 4</td>
<td></td>
</tr>
<tr>
<td>0.50</td>
<td>2</td>
<td>2.50 (\times) 2 = 5</td>
<td></td>
</tr>
<tr>
<td>0.67</td>
<td>3</td>
<td>1.67 (\times) 3 = 5</td>
<td></td>
</tr>
</tbody>
</table>
SAMPLE PROBLEM 7.10 Calculating an Empirical Formula Using Multipliers

Ascorbic acid (vitamin C), found in citrus fruits and vegetables, is important in metabolic reactions in the body, in the synthesis of collagen, and in the prevention of scurvy. If the mass percent composition of ascorbic acid is 40.9% C, 4.58% H, and 54.5% O, what is the empirical formula of ascorbic acid?

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>40.9% C, 4.58% H, 54.5% O</td>
<td>empirical formula</td>
<td>molar mass</td>
</tr>
</tbody>
</table>

STEP 1 Calculate the moles of each element. In a sample size of 100. g, there are 40.9 g of C, 4.58 g of H, and 54.5 g of O.

\[
\begin{align*}
40.9 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} &= 3.41 \text{ mol of C} \\
4.58 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} &= 4.54 \text{ mol of H} \\
54.5 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} &= 3.41 \text{ mol of O}
\end{align*}
\]

STEP 2 Divide by the smallest number of moles. For this problem, the smallest number of moles is 3.41.

\[
\begin{align*}
\frac{3.41 \text{ mol C}}{3.41} &= 1.00 \text{ mol of C} \\
\frac{4.54 \text{ mol H}}{3.41} &= 1.33 \text{ mol of H} \\
\frac{3.41 \text{ mol O}}{3.41} &= 1.00 \text{ mol of O}
\end{align*}
\]

STEP 3 Use the lowest whole-number ratio of moles as subscripts. As calculated, the ratio of moles gives the formula

\[C_{1.00}H_{1.33}O_{1.00}\]

Because the subscript for H has a decimal that is greater than 0.1 and less than 0.9, it should not be rounded off. Then, we multiply each of the subscripts by 3 to obtain a whole number for H, which is 4. Thus, the empirical formula of ascorbic acid is \(C_3H_4O_3\).

\[C_{(1.00\times3)}H_{(1.33\times3)}O_{(1.00\times3)} = C_{3.00}H_{3.99}O_{3.00} \rightarrow C_3H_4O_3\]

STUDY CHECK 7.10

Glyoxylic acid is used by plants and bacteria to convert fats into glucose. What is the empirical formula of glyoxylic acid if it has a mass percent composition of 32.5% C, 2.70% H, and 64.8% O?

ANSWER

\(C_3H_2O_3\)
**7.5 Empirical Formulas**

**LEARNING GOAL** From the mass percent composition, calculate the empirical formula for a compound.

### 7.43 Calculate the empirical formula for each of the following:

a. 3.57 g of N and 2.04 g of O

b. 7.00 g of C and 1.75 g of H

c. 0.175 g of H, 2.44 g of N, and 8.38 g of O

d. 2.06 g of Ca, 2.66 g of Cr, and 3.28 g of O

### 7.44 Calculate the empirical formula for each of the following:

a. 2.90 g of Ag and 0.430 g of S

b. 2.22 g of Na and 0.774 g of O

c. 2.11 g of Na, 0.0900 g of H, 2.94 g of S, and 5.86 g of O

d. 5.52 g of K, 1.45 g of P, and 3.00 g of O

### 7.45 Calculate the empirical formula for each of the following:

a. 70.9% K and 29.1% S

b. 55.0% Ga and 45.0% F

c. 31.0% B and 69.0% O

d. 18.8% Li, 16.3% C, and 64.9% O

e. 51.7% C, 6.95% H, and 41.3% O

### 7.46 Calculate the empirical formula for each of the following:

a. 55.5% Ca and 44.5% S

b. 78.3% Ba and 21.7% F

c. 76.0% Zn and 24.0% P

d. 29.1% Na, 40.6% S, and 30.3% O

e. 19.8% C, 2.20% H, and 78.0% Cl

**7.6 Molecular Formulas**

**LEARNING GOAL** Determine the molecular formula of a substance from the empirical formula and molar mass.

An empirical formula represents the lowest whole-number ratio of atoms in a compound. However, empirical formulas do not necessarily represent the actual number of atoms in a molecule. A molecular formula is related to the empirical formula by a small integer such as 1, 2, or 3.

\[
\text{Molecular formula} = \text{small integer} \times \text{empirical formula}
\]

For example, in **TABLE 7.4**, we see several different compounds that have the same empirical formula, CH₂O. The molecular formulas are related to the empirical formulas by small whole numbers (integers). The same relationship is true for the molar mass and empirical formula mass. The molar mass of each of the different compounds is related to the mass of the empirical formula (30.03 g) by the same small integer.

| Table 7.4 Comparing the Molar Mass of Some Compounds with the Empirical Formula of CH₂O |
|-----------------------------------------------|----------|----------------|----------------|----------------|
| Compound          | Empirical Formula | Molecular Formula | Molar Mass (g) | Integer × Empirical Formula | Integer × Empirical Formula Mass |
| Acetaldehyde      | CH₂O       | CH₂O             | 30.03          | 1(CH₂O)          | 1 × 30.03                     |
| Acetic acid       | CH₂O       | C₂H₄O₂           | 60.06          | 2(CH₂O)          | 2 × 30.03                     |
| Lactic acid       | CH₂O       | C₃H₆O₃           | 90.09          | 3(CH₂O)          | 3 × 30.03                     |
| Erythrose         | CH₂O       | C₄H₈O₄           | 120.12         | 4(CH₂O)          | 4 × 30.03                     |
| Ribose            | CH₂O       | C₅H₁₀O₅          | 150.15         | 5(CH₂O)          | 5 × 30.03                     |
Relating Empirical and Molecular Formulas

Once we determine the empirical formula, we can calculate the empirical formula mass in grams. If we are given the molar mass of the compound, we can calculate the value of the small integer.

$$\text{Small integer} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}}$$

For example, when the molar mass of ribose is divided by the empirical formula mass, the integer is 5.

$$\text{Small integer} = \frac{\text{molar mass of ribose}}{\text{empirical formula mass of CH}_2\text{O}} = \frac{150.15 \text{ g}}{30.03 \text{ g}} = 5$$

Multiplying the subscripts in the empirical formula (CH$_2$O) by 5 gives the molecular formula of ribose, C$_3$H$_{10}$O$_5$.

$$5 \times \text{empirical formula (CH}_2\text{O)} = \text{molecular formula (C}_3\text{H}_{10}\text{O}_5)$$

Calculating a Molecular Formula

Earlier, in Sample Problem 7.10, we determined that the empirical formula of ascorbic acid (vitamin C) was C$_3$H$_5$O$_3$. If the molar mass for ascorbic acid is 176.12 g, we can calculate its molecular formula as follows:

The mass of the empirical formula C$_3$H$_5$O$_3$ is obtained in the same way as molar mass.

$$\text{Empirical formula} = 3 \text{ mol of C} + 4 \text{ mol of H} + 3 \text{ mol of O}$$

$$\text{Empirical formula mass} = (3 \times 12.01 \text{ g}) + (4 \times 1.008 \text{ g}) + (3 \times 16.00 \text{ g})$$

$$= 88.06 \text{ g}$$

$$\text{Small integer} = \frac{\text{molar mass of ascorbic acid}}{\text{empirical formula mass of C}_3\text{H}_5\text{O}_3} = \frac{176.12 \text{ g}}{88.06 \text{ g}} = 2$$

Multiplying all the subscripts in the empirical formula of ascorbic acid by 2 gives its molecular formula.

$$\text{C}_{(3 \times 3)}\text{H}_{(4 \times 2)}\text{O}_{(3 \times 2)} = \text{C}_6\text{H}_6\text{O}_6 \quad \text{Molecular formula}$$

**SAMPLE PROBLEM 7.11 Determination of a Molecular Formula**

Melamine, which is used to make plastic items such as dishes and toys, contains 28.57% C, 4.80% H, and 66.64% N. If the experimental molar mass is 125 g, what is the molecular formula of melamine?

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>28.57% C, 4.80% H, 66.64% N, molar mass 125 g</td>
<td>molecular formula</td>
<td>empirical formula mass</td>
</tr>
</tbody>
</table>

**STEP 1** Obtain the empirical formula and calculate the empirical formula mass. In a sample size of 100 g of this compound, there are 28.57 g of C, 4.80 g of H, and 66.64 g of N.

$$28.57 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.38 \text{ mol of C}$$

Brightly colored dishes are made of melamine.

**Guide to Calculating a Molecular Formula from an Empirical Formula**

**STEP 1** Obtain the empirical formula and calculate the empirical formula mass.

**STEP 2** Divide the molar mass by the empirical formula mass to obtain a small integer.

**STEP 3** Multiply the empirical formula by the small integer to obtain the molecular formula.
4.80 g H × \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.76 \text{ mol of H}

66.64 g N × \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 4.76 \text{ mol of N}

Divide the moles of each element by the smallest number of moles, 2.38, to obtain the subscripts of each element in the formula.

\[
\begin{align*}
2.38 \text{ mol C} & = 1.00 \text{ mol of C} \\
4.76 \text{ mol H} & = 2.00 \text{ mol of H} \\
4.76 \text{ mol N} & = 2.00 \text{ mol of N}
\end{align*}
\]

Using these values as subscripts, \(C_{1.00}H_{2.00}N_{2.00}\), we write the empirical formula for melamine as \(CH_2N_2\).

\[
C_{1.00}H_{2.00}N_{2.00} = CH_2N_2 \quad \text{Empirical formula}
\]

Now we calculate the molar mass for this empirical formula as follows:

\[
\text{Empirical formula mass} = (1 \times 12.01) + (2 \times 1.008) + (2 \times 14.01) = 42.05 \text{ g}
\]

**STEP 2** Divide the molar mass by the empirical formula mass to obtain a small integer.

\[
\text{Small integer} = \frac{\text{molar mass of melamine}}{\text{empirical formula mass of CH}_2N_2} = \frac{125 \text{ g}}{42.05 \text{ g}} = 2.97
\]

**STEP 3** Multiply the empirical formula by the small integer to obtain the molecular formula. Because the experimental molar mass is close to 3 times the empirical formula mass, the subscripts in the empirical formula are multiplied by 3 to give the molecular formula.

\[
C_{(1\times3)}H_{(2\times3)}N_{(2\times3)} = C_3H_6N_6 \quad \text{Molecular formula}
\]

**STUDY CHECK 7.11**
The insecticide lindane has a mass percent composition of 24.78% C, 2.08% H, and 73.14% Cl. If its experimental molar mass is 290 g, what is the molecular formula?

**ANSWER**
\(C_6H_6Cl_6\)

---

**QUESTIONS AND PROBLEMS**

### 7.6 Molecular Formulas

**LEARNING GOAL** Determine the molecular formula of a substance from the empirical formula and molar mass.

**Applications**

#### 7.47 Write the empirical formula for each of the following substances:

- a. \(H_2O_2\), peroxide
- b. \(C_{18}H_{12}\), chrysene, used in the manufacture of dyes
- c. \(C_{10}H_{16}O_2\), chrysanthemic acid, in pyrethrum flowers
- d. \(C_9H_{18}N_6\), altretamine, an anticancer medication
- e. \(C_2H_4N_2O_2\), oxamide, a fertilizer

#### 7.48 Write the empirical formula for each of the following substances:

- a. \(C_6H_{10}O_5\), pyrogallol, a developer in photography
- b. \(C_6H_{12}O_6\), galactose, a carbohydrate
- c. \(C_8H_{10}O_4\), terephthalic acid, used in the manufacture of plastic bottles
- d. \(C_6Cl_{12}\), hexachlorobenzene, a fungicide
- e. \(C_24H_{16}O_{12}\), laccaic acid, a crimson dye
7.49 The carbohydrate fructose found in honey and fruits has an empirical formula of $\text{CH}_2\text{O}$. If the experimental molar mass of fructose is 180 g, what is its molecular formula?

7.50 Caffeine has an empirical formula of $\text{C}_4\text{H}_5\text{N}_2\text{O}$. If it has an experimental molar mass of 194 g, what is the molecular formula of caffeine?

7.51 Benzene and acetylene have the same empirical formula, CH. However, benzene has an experimental molar mass of 78 g, and acetylene has an experimental molar mass of 26 g. What are the molecular formulas of benzene and acetylene?

7.52 Glyoxal, used in textiles; maleic acid, used to retard oxidation of fats and oils; and aconitic acid, a plasticizer, all have the same empirical formula, CHO. However, the experimental molar masses are glyoxal 58 g, maleic acid 117 g, and aconitic acid 174 g. What are the molecular formulas of glyoxal, maleic acid, and aconitic acid?

7.53 Mevalonic acid is involved in the biosynthesis of cholesterol. Mevalonic acid contains 48.64% C, 8.16% H, and 43.20% O. If mevalonic acid has an experimental molar mass of 148 g, what is its molecular formula?

7.54 Chloral hydrate, a sedative, contains 14.52% C, 1.83% H, 64.30% Cl, and 19.35% O. If it has an experimental molar mass of 165 g, what is the molecular formula of chloral hydrate?

7.55 Vanillic acid contains 57.14% C, 4.80% H, and 38.06% O, and has an experimental molar mass of 168 g. What is the molecular formula of vanillic acid?

7.56 Lactic acid, the substance that builds up in muscles during exercise, has a mass percent composition of 40.0% C, 6.71% H, and 53.3% O, and an experimental molar mass of 90 g. What is the molecular formula of lactic acid?

7.57 A sample of nicotine, a poisonous compound found in tobacco leaves, contains 74.0% C, 8.70% H, and 17.3% N. If the experimental molar mass of nicotine is 162 g, what is its molecular formula?

7.58 Adenine, a nitrogen-containing compound found in DNA and RNA, contains 44.5% C, 3.70% H, and 51.8% N. If adenine has an experimental molar mass of 135 g, what is its molecular formula?

Follow Up

TWO PRESCRIPTIONS FOR MAX

After completing the examination and lab tests, Chris, the veterinarian, prescribed two drugs for Max: Clavamox and Proin.

Applications

7.59 Clavamox, a broad-spectrum antibiotic, contains clavulanic acid, which has a molecular formula of $\text{C}_8\text{H}_9\text{N}_0\text{O}_5$.

a. What is the molar mass of clavulanic acid?

b. What is the mass percent of C in clavulanic acid?

c. Max weighs 29 kg. If the dose of clavulanic acid is 2.5 mg/kg, how many moles of clavulanic acid were given?

7.60 Proin is used to treat urinary tract infections. The molecular formula of Proin is $\text{C}_9\text{H}_1\text{3N}_0\text{O}$.  

a. What is the molar mass of Proin?

b. What is the mass percent of N in Proin?

c. Max weighs 29 kg. If the dose of Proin is 2.0 mg/kg, how many moles of Proin were given?
**CHAPTER REVIEW**

### 7.1 The Mole

**LEARNING GOAL** Use Avogadro’s number to calculate the number of particles in a given number of moles. Calculate the number of moles of an element in a given number of moles of a compound.

- One mole of an element contains $6.022 \times 10^{23}$ atoms.
- One mole of a compound contains $6.022 \times 10^{23}$ molecules or formula units.

### 7.2 Molar Mass

**LEARNING GOAL** Given the chemical formula of a substance, calculate its molar mass.

- The molar mass (g/mol) of a substance is the mass in grams equal numerically to its atomic mass, or the sum of the atomic masses, which have been multiplied by their subscripts in a formula.

### 7.3 Calculations Using Molar Mass

**LEARNING GOAL** Given the number of moles (or grams) of a substance, calculate the grams (or moles). Given the grams of a compound, calculate the grams of one of the elements.

- The molar mass is used as a conversion factor to change a quantity from grams to moles, or from moles to grams.

### 7.4 Mass Percent Composition

**LEARNING GOAL** Given the formula of a compound, calculate the mass percent composition.

- The mass percent composition is obtained by dividing the mass in grams of each element in a compound by the mass of that compound.

### 7.5 Empirical Formulas

**LEARNING GOAL** From the mass percent composition, calculate the empirical formula for a compound.

- The empirical formula is calculated by determining the lowest whole-number mole ratio from the grams of the elements present in a sample.
- If mole ratios for an empirical formula are not all whole numbers, multiply all values by an integer to give whole numbers.

### 7.6 Molecular Formulas

**LEARNING GOAL** Determine the molecular formula of a substance from the empirical formula and molar mass.

- A molecular formula is equal to, or a multiple of, the empirical formula.
- The experimental molar mass, which must be known, is divided by the mass of the empirical formula to obtain the small integer used to convert the empirical formula to the molecular formula.
**KEY TERMS**

**Avogadro’s number** The number of items in a mole, equal to 6.022 x 10^23.

**empirical formula** The simplest or smallest whole-number ratio of the atoms in a formula.

**formula unit** The group of ions represented by the formula of an ionic compound.

**mass percent composition** The percent by mass of the elements in a formula.

**molar mass** The mass in grams of 1 mol of an element is equal numerically to its atomic mass. The molar mass of a compound is equal to the sum of the masses of the elements in the formula.

**mole** A group of atoms, molecules, or formula units that contains 6.022 x 10^23 of these items.

**molecular formula** The actual formula that gives the number of atoms of each type of element in the compound.

---

**CORE CHEMISTRY SKILLS**

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

**Converting Particles to Moles (7.1)**
- In chemistry, atoms, molecules, and ions are counted by the mole (abbreviated mol in calculations), a unit that contains 6.022 x 10^23 items, which is Avogadro’s number.
- For example, 1 mol of carbon contains 6.022 x 10^23 atoms of carbon and 1 mol of H₂O contains 6.022 x 10^23 molecules of H₂O.
- Avogadro’s number is used to convert between particles and moles.

**Example:** How many moles of nickel contain 2.45 x 10^24 Ni atoms?

**Answer:**

\[
2.45 \times 10^{24} \text{ Ni atoms} \times \frac{1 \text{ mol Ni}}{6.022 \times 10^{23} \text{ Ni atoms}} = 4.07 \text{ mol of Ni}
\]

**Calculating Molar Mass (7.2)**
- The molar mass of a compound is the sum of the molar mass of each element in its chemical formula multiplied by its subscript in the formula.

**Example:** Calculate the molar mass for pinene, C₁₀H₁₆, a component of pine tree sap.

**Answer:**

\[
10 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 120.1 \text{ g of C}
\]

\[
16 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 16.13 \text{ g of H}
\]

\[
\text{Molar mass of C₁₀H₁₆} = 136.2 \text{ g}
\]

**Using Molar Mass as a Conversion Factor (7.3)**
- Molar mass is used as a conversion factor to convert between the moles and grams of a substance.

**Example:** The frame of a bicycle contains 6500 g of aluminum. How many moles of aluminum are in the bicycle frame?

**Answer:**

\[
6500 \text{ g of Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 240 \text{ mol of Al}
\]

**Calculating Mass Percent Composition (7.4)**
- The mass of a compound contains a definite proportion by mass of its elements.
- The mass percent of an element in a compound is calculated by dividing the mass of that element by the mass of the compound.

\[
\text{Mass percent of an element} = \frac{\text{mass of an element}}{\text{mass of the compound}} \times 100\%
\]

**Example:** Dinitrogen tetroxide, N₂O₄, is used in liquid fuels for rockets. If it has a molar mass of 92.02 g, what is the mass percent of nitrogen?

**Answer:**

\[
\text{Mass } \% \text{ N} = \frac{28.02 \text{ g N}}{92.02 \text{ g N₂O₄}} \times 100\% = 30.45\% \text{ N}
\]

**Calculating an Empirical Formula (7.5)**
- The empirical formula or simplest formula represents the lowest whole-number ratio of the atoms and therefore moles of elements in a compound.
- For example, dinitrogen tetroxide, N₂O₄, has the empirical formula NO₂.
- To calculate the empirical formula, the grams of each element are converted to moles and divided by the smallest number of moles to obtain the lowest whole-number ratio.

**Example:** Calculate the empirical formula for a compound that contains 3.28 g of Cr and 6.72 g of Cl.
Answer:

- Convert the grams of each element to moles.
  
  \[
  \frac{3.28 \text{ g Cr}}{52.00 \text{ g Cr}} = 0.0631 \text{ mol of Cr}
  \]
  
  \[
  \frac{6.72 \text{ g Cl}}{35.45 \text{ g Cl}} = 0.190 \text{ mol of Cl}
  \]

Divide by the smallest number of moles (0.0631) to obtain the empirical formula.

\[
\frac{0.0631 \text{ mol Cr}}{0.0631} = 1.00 \text{ mol of Cr}
\]

\[
\frac{0.190 \text{ mol Cl}}{0.0631} = 3.01 \text{ mol of Cl}
\]

Write the empirical formula using the whole-number ratios of moles.

\[
\text{Cr}_{1.00}\text{Cl}_{3.01} \rightarrow \text{CrCl}_3
\]

Calculating a Molecular Formula (7.6)

- A molecular formula is related to the empirical formula by a small whole number (integer) such as 1, 2, or 3.
  
  Molecular formula = small integer \times \text{empirical formula}

- If the empirical formula mass and the molar mass are known for a compound, an integer can be calculated by dividing the molar mass by the empirical formula mass.
  
  \[
  \text{Small integer} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}}
  \]

Example: Cymene, a component in oil of thyme, has an empirical formula \(\text{C}_3\text{H}_7\) and an experimental molar mass of 135 g. What is the molecular formula of cymene?

Answer:

- Empirical formula = 5 mol of \(\text{C}\) + 7 mol of \(\text{H}\)
  
  \[
  \text{Empirical formula mass} = (5 \times 12.01 \text{ g}) + (7 \times 1.008 \text{ g})
  \]
  
  \[
  = 67.11 \text{ g}
  \]

  \[
  \frac{\text{Molar mass of cymene}}{\text{Empirical formula mass of } \text{C}_3\text{H}_7} = \frac{135 \text{ g}}{67.11 \text{ g}}
  \]
  
  \[
  = 2.01 \text{ (round off to 2)}
  \]

  The molecular formula of cymene is calculated by multiplying each of the subscripts in the empirical formula by 2.
  
  \[
  \text{C}_{(5 \times 2)}\text{H}_{(7 \times 2)} = \text{C}_{10}\text{H}_{14}
  \]

UNDERSTANDING THE CONCEPTS

The chapter sections to review are shown in parentheses at the end of each question.

7.61 A dandruff shampoo contains dipyrithione, an antibacterial and antifungal agent, which has the empirical formula \(\text{C}_6\text{H}_7\text{NOS}\). (7.1, 7.2, 7.3, 7.4, 7.5)

- A. What is the mass of the empirical formula of dipyrithione?
- B. If the molar mass of dipyrithione is 252.31 g, what is its molecular formula?
- C. What is the mass percent of S in dipyrithione?
- D. How many grams of H are there in 10.0 g of dipyrithione?
- E. What is the weight of 1.5 \times 10^{-3} \text{ mol of dipyrithione}?

7.62 Ibuprofen, the anti-inflammatory drug in Advil, has the formula \(\text{C}_{13}\text{H}_{10}\text{O}_2\). (7.1, 7.2, 7.3, 7.4)

- A. What is the empirical formula of ibuprofen?
- B. What is the molar mass of ibuprofen?
- C. What is the mass percent of O in ibuprofen?
- D. How many grams of C are in 0.425 g of ibuprofen?
- E. How many moles of ibuprofen are in 2.45 g of ibuprofen?

7.63 Using the models of the molecules (black = C, white = H, yellow = S, green = Cl), determine each of the following for models of compounds 1 and 2: (7.2, 7.4, 7.5, 7.6)

1. \[
\]
2. \[
\]

- A. molecular formula
- B. empirical formula
- C. molar mass
- D. mass percent composition

7.64 Using the models of the molecules (black = C, white = H, yellow = S, red = O), determine each of the following for models of compounds 1 and 2: (7.2, 7.4, 7.5, 7.6)

1. \[
\]
2. \[
\]

- A. molecular formula
- B. empirical formula
- C. molar mass
- D. mass percent composition
7.65 Calculate the molar mass for each of the following: (7.5)
   a. C\textsubscript{12}H\textsubscript{22}O\textsubscript{11}, sucrose, is commonly known as table sugar
   b. C\textsubscript{18}H\textsubscript{34}O\textsubscript{2}, lauric acid, from coconut oil
   c. FeSO\textsubscript{4}, ferrous sulfate, iron supplement
   d. Ca\textsubscript{3}(CrO\textsubscript{4})\textsubscript{2}, calcium citrate, calcium supplement

7.66 Calculate the molar mass for each of the following: (7.5)
   a. C\textsubscript{8}H\textsubscript{10}N\textsubscript{2}O\textsubscript{2}, caffeine found in coffee beans
   b. TiO\textsubscript{2}, titania dioxide used in sunscreen lotion
   c. Al(OH)\textsubscript{3}, aluminum hydroxide used to relieve heartburn and upset stomach
   d. Ca\textsubscript{3}(PO\textsubscript{4})\textsubscript{2}, calcium phosphate used to prevent and treat calcium deficiency

7.67 Calculate the mass, in grams, of N in each of the following: (7.3)
   a. 4.9 g of aspartic acid (C\textsubscript{4}H\textsubscript{7}N\textsubscript{2}O\textsubscript{4})
   b. 2.4 g of melamine (C\textsubscript{4}H\textsubscript{6}N\textsubscript{4})
   c. 9.3 g of ammonium chloride (NH\textsubscript{4}Cl)

7.68 Calculate the mass, in grams, of Cl in each of the following: (7.3)
   a. 3.2 g of chloroform (CHCl\textsubscript{3})
   b. 3.2 g of sodium hypochlorite (NaClO)
   c. 3.2 g of aluminum chloride (AlCl\textsubscript{3})

7.69 Calculate the mass percent composition for each of the following compounds: (7.4)
   a. 3.85 g of Ca and 3.65 g of F
   b. 0.389 g of Na and 0.271 g of O
   c. 12.4 g of K, 17.4 g of Mn, and 20.3 g of O

7.70 Calculate the mass percent composition for each of the following compounds: (7.4)
   a. 0.457 g of C and 0.043 g of H
   b. 3.65 g of Na, 2.54 g of S, and 3.81 g of O
   c. 0.907 g of Na, 1.40 g of Cl, and 1.89 g of O

7.71 Calculate the mass percent composition for each of the following compounds: (7.4)
   a. K\textsubscript{2}CrO\textsubscript{4}
   b. Al(HCO\textsubscript{3})\textsubscript{3}
   c. C\textsubscript{8}H\textsubscript{12}O\textsubscript{6}

7.72 Calculate the mass percent composition for each of the following compounds: (7.4)
   a. CaCO\textsubscript{3}
   b. NaC\textsubscript{2}H\textsubscript{5}O\textsubscript{2}
   c. Ba(NO\textsubscript{3})\textsubscript{2}

7.73 Aspirin, C\textsubscript{9}H\textsubscript{8}O\textsubscript{4}, is used to reduce inflammation and reduce fever. (7.1, 7.2, 7.3, 7.4)
   a. What is the mass percent composition of aspirin?
   b. How many moles of aspirin contain 5.0 \times 10\textsuperscript{22} atoms of C?
   c. How many grams of O are in 7.50 g of aspirin?
   d. How many molecules of aspirin contain 2.50 g of H?

7.74 Ammonium sulfate, (NH\textsubscript{4})\textsubscript{2}SO\textsubscript{4}, is used in fertilizers. (7.1, 7.2, 7.3, 7.4)
   a. What is the mass percent composition of (NH\textsubscript{4})\textsubscript{2}SO\textsubscript{4}?
   b. How many atoms of H are in 0.75 mol of (NH\textsubscript{4})\textsubscript{2}SO\textsubscript{4}?
   c. How many grams of O are in 4.50 \times 10\textsuperscript{23} formula units of (NH\textsubscript{4})\textsubscript{2}SO\textsubscript{4}?
   d. What mass of (NH\textsubscript{4})\textsubscript{2}SO\textsubscript{4} contains 2.50 g of S?

7.75 A mixture contains 0.250 mol of Mn\textsubscript{2}O\textsubscript{3} and 20.0 g of MnO\textsubscript{2}. (7.1, 7.2, 7.3)
   a. How many moles of O are present in the mixture?
   b. How many grams of Mn are in the mixture?

7.76 A mixture contains 4.00 \times 10\textsuperscript{23} molecules of PCl\textsubscript{3} and 0.250 mol of PCl\textsubscript{5}. (7.1, 7.2, 7.3)
   a. How many grams of Cl are present in the mixture?
   b. How many moles of P are in the mixture?
**CHALLENGE QUESTIONS**

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

7.87 A toothpaste contains 0.240% by mass sodium fluoride used to prevent tooth decay and 0.30% by mass triclosan, C_{13}H_{21}Cl_{2}O_{3}, a preservative and antingivitis agent. One tube contains 119 g of toothpaste. (7.1, 7.2, 7.3, 7.4)

Components in toothpaste include triclosan and NaF.

a. How many moles of NaF are in the tube of toothpaste?
b. How many fluoride ions, F\(^{-}\), are in the tube of toothpaste?
c. How many grams of sodium ion, Na\(^{+}\), are in 1.50 g of toothpaste?
d. How many molecules of triclosan are in the tube of toothpaste?
e. What is the mass percent composition of triclosan?

7.88 Lactic acid is formed by the breakdown of lactose in cheese. It has a mass percent composition of 40.00% C, 6.72% H, and 53.28% O. If lactic acid has an experimental molar mass of 90 g, what is its molecular formula? (7.1, 7.2, 7.3, 7.4, 7.5, 7.6)

A toothpaste contains 0.240% by mass sodium fluoride used to prevent tooth decay and 0.30% by mass triclosan, C_{13}H_{21}Cl_{2}O_{3}, a preservative and antingivitis agent. One tube contains 119 g of toothpaste. (7.1, 7.2, 7.3, 7.4)

Challenge Questions

Answers to Selected Questions and Problems

7.17 a. 342.2 g  
    b. 188.18 g  
    c. 365.5 g
7.19 a. 151.16 g  
    b. 498.4 g  
    c. 331.3 g
7.21 a. 34.5 g  
    b. 112 g  
    c. 9.80 \times 10^4 g
7.23 a. 3.03 g  
    b. 38.1 g  
    c. 9.30 g
7.25 a. 0.760 mol of Ag  
    b. 0.0240 mol of C  
    c. 0.881 mol of NH\(_3\)  
    d. 0.452 mol of CH\(_4\)  
    e. 1.53 mol of Fe\(_2\)O\(_3\)
7.27 a. 6.25 mol of He  
    b. 0.781 mol of O\(_2\)  
    c. 0.321 mol of Al(OH)\(_3\)  
    d. 0.106 mol of Ga\(_2\)S\(_3\)
7.29 a. 0.188 g of C  
    b. 235 g of C  
    c. 0.959 g of C  
    d. 48.9 g of C
7.31 a. 6.17 mol of H  
    b. 54.0 g of C  
    c. 27.8 g of C  
    d. 0.0465 g or 4.65 \times 10^{-2} g of H

7.89 Iron(III) chromate, a yellow powder used as a pigment in paints, contains 24.3% Fe, 33.9% Cr, and 41.8% O. If it has an experimental molar mass of 460 g, what are its empirical and molecular formulas? (7.4, 7.5, 7.6)

Iron(III) chromate is a yellow pigment used in paints.

A gold bar consists of gold atoms.

7.90 A gold bar is 5.50 cm long, 3.10 cm wide, and 0.300 cm thick. (7.1, 7.2, 7.3, 7.4, 7.5, 7.6)

a. If gold has a density of 19.3 g/cm\(^3\), what is the mass of the gold bar?
b. How many atoms of gold are in the bar?
c. When the same mass of gold combines with oxygen, the oxide product has a mass of 111 g. How many moles of O are combined with the gold?
d. What is the molecular formula of the oxide product if it is the same as the empirical formula?
7.3 a. 602 g
b. 74.7 g
7.35 a. 66.0 g of N₂O
b. 0.772 mol of N₂O
c. 21.6 g of N
7.37 a. 46.8% Si; 53.2% O
   b. 81.7% C; 18.3% H
c. 32.3% Na; 45.1% S; 22.6% O
d. 54.6% C; 9.12% H; 36.3% O
7.39 a. 39.01% Mg; 60.99% F
   b. 54.09% Ca; 43.18% O; 2.721% H
c. 40.00% C; 6.714% H; 53.29% O
d. 28.19% N; 8.115% H; 36.3% O
e. 71.55% C; 6.710% H; 22.6% N; 4.909% O
7.41 a. 25.94% N
   b. 26.19% N
c. 3.67% Na
   d. 33.47% N
e. 11.96% N
7.43 a. N₂O
   c. HNO₃
   d. CH₃
7.45 a. K₂S
   c. B₂O₃
   d. Li₂CO₃
e. C₄H₂O₂
7.47 a. HO
   c. C₄H₆O
   e. CH₂NO
7.49 C₆H₁₂O₆
7.51 benzene C₆H₆; acetylene C₂H₂
7.53 C₆H₁₂O₄
7.55 C₆H₂₂O₄
7.57 C₁₀H₁₄N₂
7.59 a. 199.16 g
   b. 48.24% C
c. 3.6 × 10⁻⁴ mol
7.61 a. 126 g
   b. C₁₀H₆N₂O₂S₂
c. 25.42% S
d. 0.32 g of H
e. 0.38 g
7.63 1. a. SCl₂
   b. SCl
c. 135.04 g
   d. 47.50% S; 52.55% Cl
2. a. C₄H₆
   b. CH
   c. 78.11 g
d. 92.25% C; 7.74% H
7.65 a. 342.12 g/mole
   b. 200.12 g/mole
c. 151.92 g/mole
d. 498.36 g/mole
7.67 a. 0.515 g
   b. 1.60 g
c. 0.32 g of H
d. 0.38 g
7.69 a. 51.3% Ca; 48.7% F
   b. 58.9% Na; 41.1% O
c. 24.8% K; 34.7% Mn; 40.5% O
7.71 a. 40.27% K; 26.78% Cr; 32.96% O
   b. 12.85% Al; 1.440% H; 17.16% C; 68.57% O
c. 40.00% C; 6.716% H; 53.29% O
d. 59.99% C; 4.47% H; 35.52% O
7.73 a. 0.92 mol of aspirin
   c. 2.66 g of O
   d. 1.87 × 10²³ molecules of aspirin
7.75 a. 1.210 mol of O
   b. 40.1 g of Mn
7.77 a. CHN
   c. C₃H₄N
   d. C₃H₇NO
7.79 a. SF₆
   c. P₄O₃
7.81 a. C₁₈H₃₄O₂
   b. 5.72 × 10²¹ molecules of oleic acid
7.83 The empirical formula is CH₂O; the molecular formula is C₆H₁₂O₆.
7.85 The molecular formula is C₅H₄O₄; the molar mass is 120.10 g.
7.87 a. 0.00680 mol of NaF
   b. 4.10 × 10¹⁰ F⁻ ions
c. 0.00197 g of Na⁺ ions
d. 7.4 × 10²⁰ molecules of triclosan
e. 49.76% C; 2.436% H; 36.74% Cl; 11.05% O
7.89 The empirical formula is Fe₂Cr₃O₁₂; the molecular formula is Fe₂Cr₃O₁₂.
For parts a to f, consider the loss of electrons by atoms of the element X, and a gain of electrons by atoms of the element Y. Element X is in Group 2A (2), Period 3, and Y is in Group 7A (17), Period 3. (4.2, 5.4, 5.5, 6.2, 6.3)

a. Which element is a metal, X or Y?
b. Which element is a nonmetal, X or Y?
c. What are the ionic charges of X and Y?
d. Write the electron configurations of the atoms X and Y.
e. Write the actual formula and name of the ionic compound indicated by the ions.

A bracelet of sterling silver marked 925 contains 92.5% silver by mass and 7.5% other metals. It has a volume of 25.6 cm³ and a density of 10.2 g/cm³. (2.6, 2.8, 4.4, 7.1, 7.5, 7.6)

a. What is the mass, in kilograms, of the sterling silver bracelet?
b. How many atoms of silver are in the bracelet?
c. Determine the number of protons and neutrons in each of the two stable isotopes of silver: \(^{107}\text{Ag}\) and \(^{109}\text{Ag}\).
d. When silver combines with oxygen, a compound forms that contains 93.10% Ag by mass. What are the name and the molecular formula of the oxide product if the molecular formula is the same as the empirical formula?

Oxalic acid, a compound found in plants and vegetables such as rhubarb, has a mass percent composition of 26.7% C, 2.24% H, and 71.1% O. Oxalic acid can interfere with respiration and cause kidney or bladder stones. If a large quantity of rhubarb leaves is ingested, the oxalic acid can be toxic. The lethal dose (LD₅₀) in rats for oxalic acid is 375 mg/kg. Rhubarb leaves contain about 0.5% by mass of oxalic acid. (2.7, 7.4, 7.5, 7.6)

a. What is the empirical formula of oxalic acid?
b. If oxalic acid has an experimental molar mass of 90. g, what is its molecular formula?
c. Using the LD₅₀, how many grams of oxalic acid would be toxic for a 160-lb person?
d. How many kilograms of rhubarb leaves would the person in part e need to eat to reach the toxic level of oxalic acid?

The active ingredient in Tums neutralizes excess stomach acid.

a. What is the chemical formula of calcium carbonate?
b. What is the molar mass of calcium carbonate?
c. How many moles of calcium carbonate are in one roll of Tums that contains 12 tablets?
d. If a person takes two Tums tablets, how many grams of calcium are obtained?
e. If the Daily Value (DV) for Ca⁺⁺ to maintain bone strength in older women is 1500 mg, how many Tums tablets are needed each day?
f. What is the mass percent composition of calcium carbonate?

Tamiflu (oseltamivir), C₁₆H₂₈N₂O₄, is a drug that is used to treat influenza. The preparation of Tamiflu begins with the extraction of shikimic acid from the seedpods of star anise. From 2.6 g of star anise, 0.13 g of shikimic acid can be obtained and used to produce one capsule containing 75 mg of Tamiflu. The usual adult dosage for treatment of influenza is two capsules of Tamiflu daily for 5 days. (2.7, 6.5, 7.2, 7.3, 7.4, 7.5, 7.6)
a. What is the empirical formula of Tamiflu?
b. What is the mass percent composition of Tamiflu?
c. What is the molecular formula of shikimic acid? (Black spheres are carbon, white spheres are hydrogen, and red spheres are oxygen.)
d. How many moles of shikimic acid are contained in 1.3 g of shikimic acid?
e. How many capsules containing 75 mg of Tamiflu could be produced from 154 g of star anise?
f. How many grams of C are in one dose (75 mg) of Tamiflu?
g. How many kilograms of Tamiflu would be needed to treat all the people in a city with a population of 500 000 people?

CL.12 The compound butyric acid gives rancid butter its characteristic odor. (2.7, 2.8, 6.5, 7.1, 7.2, 7.3, 7.4, 7.5)
a. What is the molecular formula of butyric acid? (Black spheres are carbon, white spheres are hydrogen, and red spheres are oxygen.)
b. What is the empirical formula of butyric acid?
c. What is the mass percent composition of butyric acid?
d. How many grams of C are in 0.850 g of butyric acid?
e. How many grams of butyric acid contain \(3.28 \times 10^{23}\) O atoms?
f. Butyric acid has a density of 0.959 g/mL at 20 °C. How many moles of butyric acid are contained in 0.565 mL of butyric acid?

**ANSWERS**

**CL.7**

a. X is a metal; elements in Group 2A (2) are metals.
b. Y is a nonmetal; elements in Group 7A (17) are nonmetals.
c. \(X^{2+}\), \(Y^-\)
d. \(X = 1s^22s^22p^63s^2\) \(Y = 1s^22s^22p^63s^23p^3\)
e. \(\text{MgCl}_2\), magnesium chloride

**CL.9**

a. \(\text{CH}_2\text{O}_2\)
b. \(\text{C}_2\text{H}_4\text{O}_4\)
c. 27 g of oxalic acid
d. 5 kg of rhubarb

**CL.11**

a. \(\text{C}_9\text{H}_{12}\text{NO}_2\)
b. 61.52% C; 9.033% H; 8.969% N; 20.49% O
c. \(\text{C}_7\text{H}_9\text{O}_3\)
d. \(7.5 \times 10^{-3}\) mol of shikimic acid
e. 59 capsules
f. 0.046 g of C
g. \(4 \times 10^2\) kg
Exercise physiologists work with athletes as well as patients who have been diagnosed with diabetes, heart disease, pulmonary disease, or other chronic disabilities or diseases. Patients who have been diagnosed with one of these diseases are often prescribed exercise as a form of treatment, and they are referred to an exercise physiologist. The exercise physiologist evaluates the patient’s overall health and then creates a customized exercise program for that individual. The program for an athlete might focus on reducing the number of injuries, whereas a program for a cardiac patient would focus on strengthening the heart muscles. The exercise physiologist also monitors the patient for improvement and determines if the exercise is helping to reduce or reverse the progression of the disease.

A pulse oximeter measures the pulse and the \( \text{O}_2 \) saturation in the blood.
8.1 Equations for Chemical Reactions

LEARNING GOAL Identify a balanced chemical equation; determine the number of atoms in the reactants and products.

Chemical reactions occur everywhere. The fuel in our cars burns with oxygen to make the car move and run the air conditioner. When we cook our food or bleach our hair, chemical reactions take place. In our bodies, chemical reactions convert food into molecules that build muscles and move them. In the leaves of trees and plants, carbon dioxide and water are converted into carbohydrates. Some chemical reactions are simple, whereas others are quite complex. However, they can all be written with chemical equations that chemists use to describe chemical reactions. In every chemical reaction, the atoms in the reacting substances, called reactants, are rearranged to give new substances called products.

A chemical change occurs when a substance is converted into one or more new substances. For example, when silver tarnishes, the shiny silver metal (Ag) reacts with sulfur (S) to become the dull, black substance we call tarnish (Ag₂S) (see FIGURE 8.1).

A chemical reaction always involves chemical change because atoms of the reacting substances form new combinations with new properties. For example, a chemical reaction takes place when a piece of iron (Fe) combines with oxygen (O₂) in the air to produce a new substance, rust (Fe₂O₃), which has a reddish-brown color. During a chemical change, new properties become visible, which are an indication that a chemical reaction has taken place (see TABLE 8.1).

<table>
<thead>
<tr>
<th>TABLE 8.1 Types of Evidence of a Chemical Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Change in color</td>
</tr>
<tr>
<td><img src="image1" alt="Fe, Fe₂O₃" /></td>
</tr>
<tr>
<td>Iron nails change color when they react with oxygen to form rust.</td>
</tr>
<tr>
<td>2. Formation of a gas (bubbles)</td>
</tr>
<tr>
<td><img src="image2" alt="Bubbles" /></td>
</tr>
<tr>
<td>Bubbles (gas) form when CaCO₃ reacts with acid.</td>
</tr>
<tr>
<td>3. Formation of a solid (precipitate)</td>
</tr>
<tr>
<td><img src="image3" alt="Solid" /></td>
</tr>
<tr>
<td>A yellow solid forms when potassium iodide is added to lead nitrate.</td>
</tr>
<tr>
<td>4. Heat (or a flame) produced or heat absorbed</td>
</tr>
<tr>
<td><img src="image4" alt="Flame" /></td>
</tr>
<tr>
<td>Methane gas burns in the air with a hot flame.</td>
</tr>
</tbody>
</table>

FIGURE 8.1 A chemical change produces new substances with new properties.

Q Why is the formation of tarnish a chemical change?
Writing a Chemical Equation

When you build a model airplane, prepare a new recipe, or mix a medication, you follow a set of directions. These directions tell you what materials to use and the products you will obtain. In chemistry, a chemical equation tells us the materials we need and the products that will form.

Suppose you work in a bicycle shop, assembling wheels and frames into bicycles. You could represent this process by a simple equation:

\[
\text{Equation: } 2 \text{ Wheels} + 1 \text{ Frame} \rightarrow 1 \text{ Bicycle}
\]

Reactants

Product

When you burn charcoal in a grill, the carbon in the charcoal combines with oxygen to form carbon dioxide. We can represent this reaction by a chemical equation.

What is the evidence for a chemical change in the reaction of carbon and oxygen to form carbon dioxide?

When you burn charcoal in a grill, the carbon in the charcoal combines with oxygen to form carbon dioxide. We can represent this reaction by a chemical equation.

\[
\text{Equation: } C(s) + O_2(g) \xrightarrow{\Delta} CO_2(g)
\]

Reactants

Product

In a chemical equation, the formulas of the reactants are written on the left of the arrow and the formulas of the products on the right. The chemical equation for burning carbon is balanced because there is one carbon atom and two oxygen atoms in both the reactants and the products. When there are two or more formulas on the same side, they are separated by plus (+) signs.

Generally, each formula in an equation is followed by an abbreviation, in parentheses, that gives the physical state of the substance: solid (s), liquid (l), or gas (g). If a substance is dissolved in water, it is in an aqueous (aq) solution. The delta sign (\(\Delta\)) indicates that heat was used to start the reaction. TABLE 8.2 summarizes some of the symbols used in equations.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Meaning</th>
</tr>
</thead>
<tbody>
<tr>
<td>+</td>
<td>Separates two or more formulas</td>
</tr>
<tr>
<td>---</td>
<td>Reacts to form products</td>
</tr>
<tr>
<td>(s)</td>
<td>Solid</td>
</tr>
<tr>
<td>(l)</td>
<td>Liquid</td>
</tr>
<tr>
<td>(g)</td>
<td>Gas</td>
</tr>
<tr>
<td>(aq)</td>
<td>Aqueous</td>
</tr>
<tr>
<td>(\Delta)</td>
<td>Reactants are heated</td>
</tr>
</tbody>
</table>
Identifying a Balanced Chemical Equation

When a chemical reaction takes place, the bonds between the atoms of the reactants are broken and new bonds are formed to give the products. All atoms are conserved, which means that atoms cannot be gained, lost, or changed into other types of atoms. Every chemical reaction must be written as a balanced equation, which shows the same number of atoms for each element in the reactants and in the products.

Now consider the balanced reaction in which hydrogen reacts with oxygen to form water written as follows:

\[ 2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g) \]

In the balanced equation, there are whole numbers called coefficients in front of the formulas. On the reactant side, the coefficient of 2 in front of the H\(_2\) formula represents two molecules of hydrogen, which is 4 atoms of H. A coefficient of 1 is understood for O\(_2\), which gives 2 atoms of O. On the product side, the coefficient of 2 in front of the H\(_2\)O formula represents 2 molecules of water. Because the coefficient of 2 multiplies all the atoms in H\(_2\)O, there are 4 hydrogen atoms and 2 oxygen atoms in the products. Because there are the same number of hydrogen atoms and oxygen atoms in the reactants as in the products, we know that the equation is balanced. This illustrates the Law of Conservation of Matter, which states that matter cannot be created or destroyed during a chemical reaction.

**SAMPLE PROBLEM 8.1 Number of Atoms in Balanced Chemical Equations**

Indicate the number of each type of atom in the following balanced chemical equation:

\[ \text{Fe}_2\text{S}_3(s) + 6\text{HCl}(aq) \rightarrow 2\text{FeCl}_3(aq) + 3\text{H}_2\text{S}(g) \]

<table>
<thead>
<tr>
<th></th>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>S</td>
<td>3</td>
<td>3</td>
</tr>
<tr>
<td>H</td>
<td>6</td>
<td>6</td>
</tr>
<tr>
<td>Cl</td>
<td>0</td>
<td>0</td>
</tr>
</tbody>
</table>

**TRY IT FIRST**

**SOLUTION**

The total number of atoms in each formula is obtained by multiplying the coefficient by each subscript in the chemical formula.

<table>
<thead>
<tr>
<th></th>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>S</td>
<td>3</td>
<td>3</td>
</tr>
<tr>
<td>H</td>
<td>6</td>
<td>6</td>
</tr>
<tr>
<td>Cl</td>
<td>0</td>
<td>0</td>
</tr>
</tbody>
</table>

**STUDY CHECK 8.1**

When ethane, C\(_2\)H\(_6\), burns in oxygen, the products are carbon dioxide and water. The balanced equation is written as

\[ 2\text{C}_2\text{H}_6(g) + 7\text{O}_2(g) \xrightarrow{\Delta} 4\text{CO}_2(g) + 6\text{H}_2\text{O}(g) \]

Calculate the number of each type of atom in the reactants and in the products.

**ANSWER**

In both the reactants and products, there are 4 C atoms, 12 H atoms, and 14 O atoms.
# QUESTIONS AND PROBLEMS

## 8.1 Equations for Chemical Reactions

**LEARNING GOAL** Identify a balanced chemical equation; determine the number of atoms in the reactants and products.

### 8.1

1. State the number of atoms of oxygen in the reactants and products for each of the following equations:
   - a. $3\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{NO}(g) + 2\text{HNO}_3(aq)$
   - b. $5\text{C}(s) + 2\text{SO}_2(g) \rightarrow \text{CS}_2(g) + 4\text{CO}(g)$
   - c. $2\text{C}_2\text{H}_2(g) + 5\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 2\text{H}_2\text{O}(g)$
   - d. $\text{N}_2\text{H}_4(g) + 2\text{H}_2\text{O}_2(g) \rightarrow \text{N}_2(g) + 4\text{H}_2\text{O}(g)$

2. State the number of atoms of oxygen in the reactants and products for each of the following equations:
   - a. $\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)$
   - b. $4\text{P}(s) + 5\text{O}_2(g) \rightarrow 2\text{P}_2\text{O}_5(s)$
   - c. $4\text{NH}_3(g) + 6\text{NO}(g) \rightarrow 5\text{N}_2(g) + 6\text{H}_2\text{O}(g)$
   - d. $6\text{CO}_2(g) + 6\text{H}_2\text{O}(l) \rightarrow \text{C}_6\text{H}_12\text{O}_6(aq) + 6\text{O}_2(g)$

3. Determine whether each of the following equations is balanced or not balanced:
   - a. $\text{S}(s) + \text{O}_2(g) \rightarrow \text{SO}_2(g)$
   - b. $\text{Al}(s) + \text{Cl}_2(g) \rightarrow \text{AlCl}_3(s)$
   - c. $2\text{NaOH}(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{Na}_2\text{SO}_4(aq)$
   - d. $\text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g)$

4. Determine whether each of the following equations is balanced or not balanced:
   - a. $\text{PCl}_3(l) + \text{Cl}_2(g) \rightarrow \text{PCl}_5(s)$
   - b. $\text{CO}(g) + 2\text{H}_2(g) \rightarrow \text{CH}_3\text{OH}(l)$
   - c. $2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + \text{O}_2(g)$
   - d. $\text{Mg}(s) + \text{N}_2(g) \rightarrow \text{Mg}_3\text{N}_2(s)$

5. All of the following are balanced equations. State the number of atoms of each element in the reactants and in the products.
   - a. $2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s)$
   - b. $\text{PCl}_3(l) + 3\text{H}_2(g) \rightarrow \text{PH}_3(g) + 3\text{HCl}(g)$
   - c. $\text{P}_4\text{O}_6(s) + 6\text{H}_2\text{O}(l) \rightarrow 4\text{H}_3\text{PO}_4(aq)$

6. All of the following are balanced equations. State the number of atoms of each element in the reactants and in the products.
   - a. $2\text{Na}(s) + 3\text{O}_2(g) \rightarrow 2\text{Na}_2\text{O}_3(s)$
   - b. $\text{Al}_2\text{O}_3(s) + 6\text{HCl}(aq) \rightarrow 3\text{H}_2\text{O}(l) + 2\text{AlCl}_3(aq)$
   - c. $\text{C}_2\text{H}_4(l) + 8\text{O}_2(g) \rightarrow 5\text{CO}_2(g) + 6\text{H}_2\text{O}(g)$

## 8.2 Balancing a Chemical Equation

**LEARNING GOAL** Write a balanced chemical equation from the formulas of the reactants and products for a chemical reaction.

The chemical reaction that occurs in the flame of a gas burner you use in the laboratory or a gas cooktop is the reaction of methane gas, CH$_4$, and oxygen to produce carbon dioxide and water. We now show the process of writing and balancing the chemical equation for this reaction in Sample Problem 8.2.

### Core Chemistry Skill: Balancing a Chemical Equation

#### Guide to Writing and Balancing a Chemical Equation

**STEP 1**
Write an equation using the correct formulas for the reactants and products.

**STEP 2**
Count the atoms of each element in the reactants and products.

**STEP 3**
Use coefficients to balance each element.

**STEP 4**
Check the final equation to confirm it is balanced.

### Sample Problem 8.2: Writing and Balancing a Chemical Equation

The reaction of methane gas (CH$_4$) and oxygen gas (O$_2$) produces the gases carbon dioxide (CO$_2$) and water (H$_2$O). Write a balanced chemical equation for this reaction.

#### Try It First

**SOLUTION**

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>reactants, products</td>
<td>balanced equation</td>
<td>equal numbers of atoms in reactants and products</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1**
Write an equation using the correct formulas for the reactants and products.

$\text{CH}_4(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$
**STEP 2** Count the atoms of each element in the reactants and products. When we compare the atoms on the reactant side and the atoms on the product side, we see that there are more H atoms in the reactants and more O atoms in the products.

\[ \text{CH}_4(g) + \text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + \text{H}_2\text{O}(g) \]

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 C atom</td>
<td>1 C atom</td>
</tr>
<tr>
<td>4 H atoms</td>
<td>2 H atoms</td>
</tr>
<tr>
<td>2 O atoms</td>
<td>3 O atoms</td>
</tr>
</tbody>
</table>

**STEP 3** Use coefficients to balance each element. We will start by balancing the H atoms in CH\(_4\) because it has the most atoms. By placing a coefficient of 2 in front of the formula for H\(_2\)O, a total of 4 H atoms in the products is obtained. *Only use coefficients to balance an equation. Do not change any of the subscripts: This would alter the chemical formula of a reactant or product.*

\[ \text{CH}_4(g) + 2\text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + 2\text{H}_2\text{O}(g) \]

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 C atom</td>
<td>1 C atom</td>
</tr>
<tr>
<td>4 H atoms</td>
<td>4 H atoms</td>
</tr>
<tr>
<td>2 O atoms</td>
<td>4 O atoms</td>
</tr>
</tbody>
</table>

We can balance the O atoms on the reactant side by placing a coefficient of 2 in front of the formula \(\text{O}_2\). There are now 4 O atoms in both the reactants and products.

\[ \text{CH}_4(g) + 2\text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + 2\text{H}_2\text{O}(g) \text{ Balanced} \]

**STEP 4** Check the final equation to confirm it is balanced. In the final equation, the numbers of atoms of C, H, and O are the same in both the reactants and the products. The equation is balanced.

\[ \text{CH}_4(g) + 2\text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + 2\text{H}_2\text{O}(g) \]

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 C atom</td>
<td>1 C atom</td>
</tr>
<tr>
<td>4 H atoms</td>
<td>4 H atoms</td>
</tr>
<tr>
<td>4 O atoms</td>
<td>4 O atoms</td>
</tr>
</tbody>
</table>

How do you check that a chemical equation is balanced?
In a balanced chemical equation, the coefficients must be the lowest possible whole numbers. Suppose you had obtained the following for the balanced equation:

$$2\text{CH}_4(g) + 4\text{O}_2(g) \xrightarrow{\Delta} 2\text{CO}_2(g) + 4\text{H}_2\text{O}(g)$$ Incorrect

Although there are equal numbers of atoms on both sides of the equation, this is not written correctly. To obtain coefficients that are the lowest whole numbers, we divide all the coefficients by 2.

**STUDY CHECK 8.2**
Balance the following chemical equation:

$$\text{Al}(s) + \text{Cl}_2(g) \rightarrow \text{AlCl}_3(s)$$

**ANSWER**

$$2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s)$$

**Whole-Number Coefficients**

Sometimes the coefficients of the compounds in an equation need to be increased to give whole numbers for the coefficients. Then we need to adjust the coefficients and count the total atoms on both sides once again as shown in Sample Problem 8.3.

**SAMPLE PROBLEM 8.3 Balancing a Chemical Equation with Whole-Number Coefficients**

Acetylene, $\text{C}_2\text{H}_2$, is used to produce high temperatures for welding by reacting it with $\text{O}_2$ to produce $\text{CO}_2$ and $\text{H}_2\text{O}$. All of the compounds are gases. Write a balanced chemical equation with whole-number coefficients for this reaction.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>reactants, products</td>
<td>balanced equation</td>
<td>equal numbers of atoms in reactants and products</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Write an equation using the correct formulas of the reactants and products.

$$\text{C}_2\text{H}_2(g) + \text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + \text{H}_2\text{O}(g)$$ Not balanced

Acetylene

**STEP 2** Count the atoms of each element in the reactants and products. When we compare the atoms on the reactant side and the product side, we see that there are more C atoms in the reactants and more O atoms in the products.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 C atoms</td>
<td>1 C atom</td>
</tr>
<tr>
<td>2 H atoms</td>
<td>2 H atoms</td>
</tr>
<tr>
<td>2 O atoms</td>
<td>3 O atoms</td>
</tr>
</tbody>
</table>

**STEP 3** Use coefficients to balance each element. We start by balancing the C in $\text{C}_2\text{H}_2$ by placing a coefficient of 2 in front of the formula for $\text{CO}_2$. When we recheck all the atoms, there are 2 C atoms and 2 H atoms in both the reactants and products.
C₃H₂(g) + O₂(g) \rightarrow Δ 2CO₂(g) + H₂O(g)

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 C atoms</td>
<td>2 C atoms</td>
</tr>
<tr>
<td>2 H atoms</td>
<td>2 H atoms</td>
</tr>
<tr>
<td>2 O atoms</td>
<td>5 O atoms</td>
</tr>
</tbody>
</table>

To balance the O atoms, we place a coefficient of \(\frac{5}{2}\) in front of the formula for O₂, which gives a total of 5 O atoms on each side of the equation.

C₃H₂(g) + \(\frac{5}{2}\)O₂(g) \rightarrow Δ 2CO₂(g) + H₂O(g)

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 C atoms</td>
<td>2 C atoms</td>
</tr>
<tr>
<td>2 H atoms</td>
<td>2 H atoms</td>
</tr>
<tr>
<td>5 O atoms</td>
<td>5 O atoms</td>
</tr>
</tbody>
</table>

Now the equation is balanced for atoms, but the coefficient for O₂ is a fraction. To obtain all whole-number coefficients, we multiply all the coefficients by 2.

2C₃H₂(g) + 5O₂(g) \rightarrow Δ 4CO₂(g) + 2H₂O(g)

**STEP 4** Check the final equation to confirm it is balanced. A check of the total number of atoms indicates that the equation is now balanced with whole-number coefficients.

2C₃H₂(g) + 5O₂(g) \rightarrow Δ 4CO₂(g) + 2H₂O(g)

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>4 C atoms</td>
<td>4 C atoms</td>
</tr>
<tr>
<td>4 H atoms</td>
<td>4 H atoms</td>
</tr>
<tr>
<td>10 O atoms</td>
<td>10 O atoms</td>
</tr>
</tbody>
</table>

**STUDY CHECK 8.3**
Balance the following chemical equation:

NO₂(g) + H₂(g) \rightarrow NH₃(g) + H₂O(g)

**ANSWER**

2NO₂(g) + 7H₂(g) \rightarrow 2NH₃(g) + 4H₂O(g)

**Equations with Polyatomic Ions**
Sometimes an equation contains the same polyatomic ion in both the reactants and the products. Then we can balance the polyatomic ions as a group on both sides of the equation as shown in Sample Problem 8.4.

**SAMPLE PROBLEM 8.4** Balancing Chemical Equations with Polyatomic Ions
Balance the following chemical equation:

Na₃PO₄(aq) + MgCl₂(aq) \rightarrow Mg₃(PO₄)₂(s) + NaCl(aq)

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>reactants, products</td>
<td>balanced equation</td>
<td>equal numbers of atoms in reactants and products</td>
<td></td>
</tr>
</tbody>
</table>
**ENGAGE**
What is the evidence for a chemical reaction when Na₃PO₄(aq) and MgCl₂(aq) are mixed?

**STEP 1** Write an equation using the correct formulas for the reactants and products.

\[
\text{Na}_3\text{PO}_4(aq) + \text{MgCl}_2(aq) \rightarrow \text{Mg}_3(\text{PO}_4)_2(s) + \text{NaCl}(aq)
\]

**STEP 2** Count the atoms of each element in the reactants and products. When we compare the number of ions in the reactants and products, we find that the equation is not balanced. In this equation, we can balance the phosphate ion as a group of atoms because it appears on both sides of the equation.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>3 Na⁺</td>
<td>1 Na⁺</td>
</tr>
<tr>
<td>1 PO₄³⁻</td>
<td>2 PO₄³⁻</td>
</tr>
<tr>
<td>1 Mg²⁺</td>
<td>3 Mg²⁺</td>
</tr>
<tr>
<td>2 Cl⁻</td>
<td>1 Cl⁻</td>
</tr>
</tbody>
</table>

**STEP 3** Use coefficients to balance each element. We begin with the formula that has the highest subscript values, which in this equation is Mg₃(PO₄)₂. The subscript 3 in Mg₃(PO₄)₂ is used as a coefficient for MgCl₂ to balance magnesium. The subscript 2 in Mg₃(PO₄)₂ is used as a coefficient for Na₃PO₄ to balance the phosphate ion.

\[
2\text{Na}_3\text{PO}_4(aq) + 3\text{MgCl}_2(aq) \rightarrow \text{Mg}_3(\text{PO}_4)_2(s) + \text{NaCl}(aq)
\]

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>6 Na⁺</td>
<td>1 Na⁺</td>
</tr>
<tr>
<td>2 PO₄³⁻</td>
<td>2 PO₄³⁻</td>
</tr>
<tr>
<td>3 Mg²⁺</td>
<td>3 Mg²⁺</td>
</tr>
<tr>
<td>6 Cl⁻</td>
<td>1 Cl⁻</td>
</tr>
</tbody>
</table>

In the reactants and products, we see that the sodium and chloride ions are not yet balanced. A coefficient of 6 is placed in front of the NaCl to balance the equation.

\[
2\text{Na}_3\text{PO}_4(aq) + 3\text{MgCl}_2(aq) \rightarrow \text{Mg}_3(\text{PO}_4)_2(s) + 6\text{NaCl}(aq)
\]

**STEP 4** Check the final equation to confirm it is balanced. A check of the total number of ions confirms the equation is balanced. A coefficient of 1 is understood and not usually written.
2Na₃PO₄(aq) + 3MgCl₂(aq) → Mg₃(PO₄)₂(s) + 6NaCl(aq) Balanced

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>6 Na⁺</td>
<td>6 Na⁺</td>
</tr>
<tr>
<td>2 PO₄³⁻</td>
<td>2 PO₄³⁻</td>
</tr>
<tr>
<td>3 Mg²⁺</td>
<td>6 Cl⁻</td>
</tr>
</tbody>
</table>

**STUDY CHECK 8.4**
Balance the following chemical equation:

Pb(NO₃)₂(aq) + AlBr₃(aq) → PbBr₂(s) + Al(NO₃)₃(aq)

**ANSWER**
3Pb(NO₃)₂(aq) + 2AlBr₃(aq) → 3PbBr₂(s) + 2Al(NO₃)₃(aq)

**QUESTIONS AND PROBLEMS**

**8.2 Balancing a Chemical Equation**

**LEARNING GOAL** Write a balanced chemical equation from the formulas of the reactants and products for a chemical reaction.

**8.7** Balance each of the following chemical equations:

a. N₂(g) + O₂(g) → NO(g)

b. HgO(s) → Hg(l) + O₂(g)

c. Fe(s) + O₂(g) → Fe₂O₃(s)

d. Na(s) + Cl₂(g) → NaCl(s)

e. Cu₂O(s) + O₂(g) → CuO(s)

**8.8** Balance each of the following chemical equations:

a. Ca(s) + Br₂(l) → CaBr₂(s)

b. P₄(s) + O₂(g) → P₂O₅(s)

c. C₃H₆(g) + O₂(g) → CO₂(g) + H₂O(g)

d. Ca(OH)₂(aq) + HNO₃(aq) → H₂O(l) + Ca(NO₃)₂(aq)

e. Fe₂O₃(s) + C(s) → Fe(s) + CO(g)

**8.9** Balance each of the following chemical equations:

a. Mg(s) + AgNO₃(aq) → Ag(s) + Mg(NO₃)₂(aq)

b. CuCO₃(s) → CuO(s) + CO₂(g)

c. Al(s) + CuSO₄(aq) → Cu(s) + Al₂(SO₄)₃(aq)

d. Pb(NO₃)₂(aq) + NaCl(aq) → PbCl₂(s) + NaNO₃(aq)

e. Al(s) + HCl(aq) → H₂(g) + AlCl₃(aq)

**8.10** Balance each of the following chemical equations:

a. Zn(s) + HNO₃(aq) → H₂(g) + Zn(NO₃)₂(aq)

b. Al(s) + H₂SO₄(aq) → H₂(g) + Al₂(SO₄)₃(aq)

c. K₂SO₄(aq) + BaCl₂(aq) → BaSO₄(s) + KCl(aq)

d. CaCO₃(s) → CaO(s) + CO₂(g)

e. AlCl₃(aq) + KOH(aq) → Al(OH)₃(s) + KCl(aq)

**8.11** Balance each of the following chemical equations:

a. Fe₂O₃(s) + CO(g) → Fe(s) + CO₂(g)

b. Li₃N(s) → Li(s) + N₂(g)

c. Al(s) + HBr(aq) → H₂(g) + AlBr₃(aq)

d. Ba(OH)₂(aq) + Na₃PO₄(aq) → Ba₃(PO₄)₂(s) + NaOH(aq)

e. As₂S₆(s) + O₂(g) → As₂O₃(s) + SO₂(g)

**8.12** Balance each of the following chemical equations:

a. K(s) + H₂O(l) → H₂(g) + KOH(aq)

b. Cr(s) + S₈(s) → Cr₂S₃(s)

c. BC₁₃(s) + H₂O(l) → H₂BO₃(aq) + HCl(aq)

d. Fe(OH)₃(s) + H₂SO₄(aq) → H₂O(l) + Fe₂(SO₄)₃(aq)

e. BaCl₂(aq) + Na₃PO₄(aq) → Ba₃(PO₄)₂(s) + NaCl(aq)

**8.13** Write a balanced equation using the correct formulas and include conditions (s, l, g, or aq) for each of the following chemical reactions:

a. Lithium metal reacts with liquid water to form hydrogen gas and aqueous lithium hydroxide.

b. Solid phosphorus reacts with chlorine gas to form solid phosphorus pentachloride.

c. Solid iron(II) oxide reacts with carbon monoxide gas to form solid iron and carbon dioxide gas.

d. Liquid pentene (C₅H₁₀) burns in oxygen gas to form carbon dioxide gas and water vapor.

e. Hydrogen sulfide gas and solid iron(III) chloride react to form solid iron(III) sulfide and hydrogen chloride gas.

**8.14** Write a balanced equation using the correct formulas and include conditions (s, l, g, or aq) for each of the following chemical reactions:

a. Solid sodium carbonate decomposes to produce solid sodium oxide and carbon dioxide gas.

b. Nitrogen oxide gas reacts with carbon monoxide gas to produce nitrogen gas and carbon dioxide gas.

c. Iron metal reacts with solid sulfur to produce solid iron(III) sulfide.

d. Solid calcium reacts with nitrogen gas to produce solid calcium nitride.

e. In the *Apollo* lunar module, hydrazine gas, N₂H₄, reacts with dinitrogen tetroxide gas to produce gaseous nitrogen and water vapor.
8.3 Types of Chemical Reactions

LEARNING GOAL. Identify a chemical reaction as a combination, decomposition, single replacement, double replacement, or combustion.

A great number of chemical reactions occur in nature, in biological systems, and in the laboratory. However, there are some general patterns among all reactions that help us classify reactions. Some reactions may fit into more than one reaction type.

Combination Reactions

In a combination reaction, two or more elements or compounds bond to form one product. For example, sulfur and oxygen combine to form the product sulfur dioxide.

In FIGURE 8.2, the elements magnesium and oxygen combine to form a single product, which is the ionic compound magnesium oxide formed from $\text{Mg}^{2+}$ and $\text{O}^{2-}$ ions.

$$2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s)$$

FIGURE 8.2 ▶ In a combination reaction, two or more substances combine to form one substance as product.

What happens to the atoms in the reactants in a combination reaction?
In other examples of combination reactions, elements or compounds combine to form a single product:

\[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g}) \]

Ammonia

\[ \text{Cu}(s) + \text{S}(s) \rightarrow \text{CuS}(s) \]

\[ \text{MgO}(s) + \text{CO}_2(\text{g}) \rightarrow \text{MgCO}_3(s) \]

**Decomposition Reactions**

In a decomposition reaction, a reactant splits into two or more simpler products. For example, when mercury(II) oxide is heated, the compound breaks apart into mercury atoms and oxygen (see FIGURE 8.3).

\[ 2\text{HgO}(s) \xrightarrow{\Delta} 2\text{Hg}(l) + \text{O}_2(\text{g}) \]

In another example of a decomposition reaction, when calcium carbonate is heated, it breaks apart into simpler compounds of calcium oxide and carbon dioxide.

\[ \text{CaCO}_3(s) \xrightarrow{\Delta} \text{CaO}(s) + \text{CO}_2(\text{g}) \]

**Replacement Reactions**

In a replacement reaction, elements in a compound are replaced by other elements. In a single replacement reaction, a reacting element switches place with an element in the other reacting compound.

In the single replacement reaction shown in FIGURE 8.4, zinc replaces hydrogen in hydrochloric acid, \( \text{HCl}(aq) \).

\[ \text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{H}_2(\text{g}) + \text{ZnCl}_2(aq) \]

In another single replacement reaction, chlorine replaces bromine in the compound potassium bromide.

\[ \text{Cl}_2(\text{g}) + 2\text{KBr}(s) \rightarrow 2\text{KCl}(s) + \text{Br}_2(\ell) \]
In a double replacement reaction, the positive ions in the reacting compounds switch places. In the reaction shown in Figure 8.5, barium ions change places with sodium ions in the reactants to form sodium chloride and a white solid precipitate of barium sulfate. The formulas of the products depend on the charges of the ions.

\[
\text{BaCl}_2(aq) + \text{Na}_2\text{SO}_4(aq) \rightleftharpoons \text{BaSO}_4(s) + 2\text{NaCl}(aq)
\]

When sodium hydroxide and hydrochloric acid (HCl) react, sodium and hydrogen ions switch places, forming water and sodium chloride.

\[
\text{NaOH}(aq) + \text{HCl}(aq) \rightleftharpoons \text{H}_2\text{O}(l) + \text{NaCl}(aq)
\]
Combustion Reactions

The burning of a candle and the burning of fuel in the engine of a car are examples of combus-
tion reactions. In a combustion reaction, a carbon-containing compound, usually a fuel, burns in oxygen from the air to produce carbon dioxide (CO₂), water (H₂O), and energy in the form of heat or a flame. For example, methane gas (CH₄) undergoes combustion when used to cook our food on a gas cooktop and to heat our homes. In the equation for the combustion of methane, each element in the fuel (CH₄) forms a compound with oxygen.

\[ \text{CH}_4(g) + 2\text{O}_2(g) \xrightarrow{\text{heat}} \text{CO}_2(g) + 2\text{H}_2\text{O}(g) + \text{energy} \]

Methane

The balanced equation for the combustion of propane (C₃H₈) is

\[ \text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \xrightarrow{\Delta} 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g) + \text{energy} \]

Propane is the fuel used in portable heaters and gas barbecues. Gasoline, a mixture of liq-
uid hydrocarbons, is the fuel that powers our cars, lawn mowers, and snow blowers.

**TABLE 8.3** summarizes the reaction types and gives examples.

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Combination</td>
<td>A + B \rightarrow AB</td>
</tr>
<tr>
<td></td>
<td>Ca(s) + Cl₂(g) \rightarrow CaCl₂(s)</td>
</tr>
<tr>
<td>Decomposition</td>
<td>AB \rightarrow A + B</td>
</tr>
<tr>
<td></td>
<td>Fe₂S₃(s) \rightarrow 2Fe(s) + 3S(s)</td>
</tr>
<tr>
<td>Single Replacement</td>
<td>A + BC \rightarrow AC + B</td>
</tr>
<tr>
<td></td>
<td>Cu(s) + 2AgNO₃(aq) \rightarrow 2Ag(s) + Cu(NO₃)₂(aq)</td>
</tr>
<tr>
<td>Double Replacement</td>
<td>AB + CD \rightarrow AD + CB</td>
</tr>
<tr>
<td></td>
<td>BaCl₂(aq) + K₂SO₄(aq) \rightarrow BaSO₄(s) + 2KCl(aq)</td>
</tr>
<tr>
<td>Combustion</td>
<td>CₓHᵧ + ZO₂(g) \xrightarrow{\Delta} XCO₂(g) + \frac{Y}{2}\text{H}_2\text{O}(g) + energy</td>
</tr>
<tr>
<td></td>
<td>CH₄(g) + 2O₂(g) \xrightarrow{\Delta} CO₂(g) + 2H₂O(g) + energy</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 8.5** Identifying Reactions

Classify each of the following as a combination, decomposition, single replacement, double replacement, or combustion reaction:

**a.** \(2\text{Fe}_3\text{O}_4(s) + 3\text{C}(s) \rightarrow 3\text{CO}_2(g) + 4\text{Fe}(s)\)

**b.** \(2\text{KClO}_3(s) \xrightarrow{\text{heat}} 2\text{KCl}(s) + 3\text{O}_2(g)\)

**c.** \(\text{C}_3\text{H}_8(g) + 3\text{O}_2(g) \xrightarrow{\text{heat}} 2\text{CO}_2(g) + 2\text{H}_2\text{O}(g) + \text{energy}\)

**TRY IT FIRST**

**SOLUTION**

**a.** In this single replacement reaction, a C atom replaces Fe in \(\text{Fe}_2\text{O}_3\) to form the com-

**b.** When one reactant breaks down to produce two products, the reaction is decomposi-

**c.** The reaction of a carbon compound with oxygen to produce carbon dioxide, water, and

energy makes this a combustion reaction.
Nitrogen gas (N\textsubscript{2}) and oxygen gas (O\textsubscript{2}) react to form nitrogen dioxide gas. Write the balanced chemical equation using the correct chemical formulas of the reactants and products, and identify the reaction type.

**ANSWER**

\[ \text{N}_2(g) + 2\text{O}_2(g) \rightarrow 2\text{NO}_2(g) \]  
Combination

---

**CHEMISTRY LINK TO HEALTH**

**Incomplete Combustion: Toxicity of Carbon Monoxide**

When a propane heater, fireplace, or woodstove is used in a closed room, there must be adequate ventilation. If the supply of oxygen is limited, incomplete combustion from burning gas, oil, or wood produces carbon monoxide. The incomplete combustion of methane in natural gas is written:

\[ 2\text{CH}_4(g) + 3\text{O}_2(g) \stackrel{\Delta}{\rightarrow} 2\text{CO}(g) + 4\text{H}_2\text{O}(g) + \text{energy} \]

**Limited oxygen supply**  
**Carbon monoxide**

Carbon monoxide (CO) is a colorless, odorless, poisonous gas. When inhaled, CO passes into the bloodstream, where it attaches to hemoglobin, which reduces the amount of oxygen (O\textsubscript{2}) reaching the cells. As a result, a person can experience a reduction in exercise capability, visual perception, and manual dexterity.

Hemoglobin is the protein that transports O\textsubscript{2} in the blood. When the amount of hemoglobin bound to CO (COH\textsubscript{b}) is about 10%, a person may experience shortness of breath, mild headache, and drowsiness. Heavy smokers can have levels of COH\textsubscript{b} in their blood as high as 9%. When as much as 30% of the hemoglobin is bound to CO, a person may experience more severe symptoms, including dizziness, mental confusion, severe headache, and nausea. If 50% or more of the hemoglobin is bound to CO, a person could become unconscious and die if not treated immediately with oxygen.

---

**SAMPLE PROBLEM 8.6 Writing an Equation for Combustion**

A portable burner is fueled with butane, C\textsubscript{4}H\textsubscript{10}. Write the reactants and products for the complete combustion of butane, and balance the equation.

**TRY IT FIRST**

**SOLUTION**

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>C\textsubscript{4}H\textsubscript{10}</td>
<td>balanced equation for combustion</td>
<td>reactants: carbon compound + O\textsubscript{2}, products: CO\textsubscript{2} + H\textsubscript{2}O</td>
</tr>
</tbody>
</table>

In a combustion reaction, butane gas reacts with the gas O\textsubscript{2} to form the gases CO\textsubscript{2}, H\textsubscript{2}O, and energy. We write the unbalanced equation as:

\[ \text{C}_4\text{H}_{10}(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) + \text{energy} \]

We can begin by using the subscripts in C\textsubscript{4}H\textsubscript{10} to balance the C atoms in CO\textsubscript{2} and the H atoms in H\textsubscript{2}O. However, this gives a total of 13 O atoms in the products. This is balanced by placing a coefficient of 13/2 in front of the formula for O\textsubscript{2}.

\[ \text{C}_4\text{H}_{10}(g) + \frac{13}{2}\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 5\text{H}_2\text{O}(g) + \text{energy} \]

13 O atoms

To obtain a whole number coefficient for O\textsubscript{2}, we need to multiply all the coefficients by 2.

\[ 2\text{C}_4\text{H}_{10}(g) + 13\text{O}_2(g) \rightarrow 8\text{CO}_2(g) + 10\text{H}_2\text{O}(g) + \text{energy} \]

**STUDY CHECK 8.6**

Ethene, used to ripen fruit, has the formula C\textsubscript{2}H\textsubscript{4}. Write a balanced chemical equation for the complete combustion of ethene.

**ANSWER**

\[ \text{C}_2\text{H}_4(g) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 2\text{H}_2\text{O}(g) + \text{energy} \]
8.3 Types of Chemical Reactions

LEARNING GOAL Identify a chemical reaction as a combination, decomposition, single replacement, double replacement, or combustion.

8.19 Classify each of the following as a combination, decomposition, single replacement, double replacement, or combustion reaction:

a. \( 2\text{Al}_2\text{O}_3(s) \xrightarrow{\triangle} 4\text{Al}(s) + 3\text{O}_2(g) \)

b. \( \text{Br}_2(l) + \text{Ba}_2^+(aq) \rightarrow \text{BaBr}_2(s) + \text{I}_2(g) \)

c. \( 2\text{C}_2\text{H}_2(g) + 5\text{O}_2(g) \xrightarrow{\triangle} 4\text{CO}_2(g) + 2\text{H}_2\text{O}(g) \)

d. \( \text{BaCl}_2(aq) + \text{K}_2\text{CO}_3(aq) \rightarrow \text{BaCO}_3(s) + 2\text{KCl}(aq) \)

e. \( \text{Pb}(s) + \text{O}_2(g) \rightarrow \text{PbO}_2(s) \)

8.20 Classify each of the following as a combination, decomposition, single replacement, double replacement, or combustion reaction:

a. \( \text{H}_2(g) + \text{Br}_2(l) \rightarrow 2\text{HBr}(g) \)

b. \( \text{AgNO}_3(aq) + \text{NaCl}(aq) \rightarrow \text{AgCl}(s) + \text{NaNO}_3(aq) \)

c. \( 2\text{H}_2\text{O}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g) \)

d. \( \text{Zn}(s) + \text{CuCl}_2(aq) \rightarrow \text{Cu}(s) + \text{ZnCl}_2(aq) \)

e. \( \text{C}_6\text{H}_6(g) + 7\text{O}_2(g) \rightarrow 5\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \)

8.21 Classify each of the following as a combination, decomposition, single replacement, double replacement, or combustion reaction:

a. \( 4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \) Fe is oxidized

b. \( \text{Mg}(s) + 2\text{AgNO}_3(aq) \rightarrow 2\text{Ag}(s) + \text{Mg(NO}_3)_2(aq) \)

c. \( \text{CuCO}_3(s) \xrightarrow{\triangle} \text{CuO}(s) + \text{CO}_2(g) \)

d. \( \text{Al}_2(\text{SO}_4)_3(aq) + 6\text{KOH}(aq) \rightarrow 2\text{Al(OH)}_3(s) + 3\text{K}_2\text{SO}_4(aq) \)

e. \( \text{C}_6\text{H}_6(g) + 6\text{O}_2(g) \xrightarrow{\triangle} 6\text{CO}_2(g) + 6\text{H}_2\text{O}(g) \)

8.22 Classify each of the following as a combination, decomposition, single replacement, double replacement, or combustion reaction:

a. \( \text{CuO}(s) + 2\text{HCl}(aq) \rightarrow \text{CuCl}_2(aq) + \text{H}_2\text{O}(l) \)

b. \( 2\text{Al}(s) + 3\text{Br}_2(l) \rightarrow 2\text{AlBr}_3(s) \)

c. \( \text{C}_6\text{H}_12(l) + 9\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 6\text{H}_2\text{O}(g) \)

d. \( \text{Fe}_2\text{O}_3(s) + 3\text{C}(s) \rightarrow 2\text{Fe}(s) + 3\text{CO}(g) \)

e. \( \text{C}_6\text{H}_12\text{O}_6(aq) \rightarrow 2\text{C}_2\text{H}_4\text{O}_2(aq) + 2\text{CO}_2(g) \)

8.23 Using Table 8.3, predict the products that would result from each of the following reactions and balance:

a. combination: \( \text{Mg}(s) + \text{Cl}_2(g) \rightarrow \)

b. decomposition: \( \text{HBr}(g) \rightarrow \)

c. single replacement: \( \text{Mg}(s) + \text{Zn(NO}_3)_2(aq) \rightarrow \)

d. double replacement: \( \text{K}_2\text{S}(aq) + \text{Pb(NO}_3)_2(aq) \rightarrow \)

e. combustion: \( \text{C}_2\text{H}_6(g) + \text{O}_2(g) \xrightarrow{\triangle} \)

8.24 Using Table 8.3, predict the products that would result from each of the following reactions and balance:

a. combination: \( \text{Cu}(s) + \text{O}_2(g) \rightarrow \)

b. combustion: \( \text{Cu}_2\text{O}(g) + \text{O}_2(g) \rightarrow \)

c. decomposition: \( \text{Pb}_2\text{O}_3(s) \rightarrow \)

d. single replacement: \( \text{KI}(s) + \text{Cl}_2(g) \rightarrow \)

e. double replacement: \( \text{CuCl}_2(aq) + \text{Na}_2\text{S}(aq) \rightarrow \)

8.4 Oxidation–Reduction Reactions

LEARNING GOAL Define the terms oxidation and reduction; identify the reactants oxidized and reduced.

Perhaps you have never heard of an oxidation and reduction reaction. However, this type of reaction has many important applications in your everyday life. When you see a rusty nail, tarnish on a silver spoon, or corrosion on metal, you are observing oxidation.

\[
4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \quad \text{Fe is oxidized}
\]

Rust

When we turn the lights on in our automobiles, an oxidation–reduction reaction within the car battery provides the electricity. On a cold, wintry day, we might build a fire. As the wood burns, oxygen combines with carbon and hydrogen to produce carbon dioxide, water, and heat. In the previous section, we called this a combustion reaction, but it is also an oxidation–reduction reaction. When we eat foods with starches in them, the starches break down to give glucose, which is oxidized in our cells to give us energy along with carbon dioxide and water. Every breath we take provides oxygen to carry out oxidation in our cells.

\[
\text{C}_6\text{H}_12\text{O}_6(aq) + 6\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 6\text{H}_2\text{O}(l) + \text{energy}
\]

Glucose

Oxidation–Reduction Reactions

In an oxidation–reduction reaction (redox), electrons are transferred from one substance to another. If one substance loses electrons, another substance must gain electrons. Oxidation is defined as the loss of electrons; reduction is the gain of electrons.
One way to remember these definitions is to use the following acronym:

**OIL RIG**
- **Oxidation** Is Loss of electrons
- **Reduction** Is Gain of electrons

**Oxidation–Reduction**

In general, atoms of metals lose electrons to form positive ions, whereas nonmetals gain electrons to form negative ions. Now we can say that metals are oxidized and nonmetals are reduced.

The green color that appears on copper surfaces from weathering, known as *patina*, is a mixture of CuCO₃ and CuO. We can now look at the oxidation and reduction reactions that take place when copper metal reacts with oxygen in the air to produce copper(II) oxide.

\[ 2\text{Cu}(s) + \text{O}_2(g) \rightarrow 2\text{CuO}(s) \]

The element Cu in the reactants has a charge of 0, but in the CuO product, it is present as \( \text{Cu}^{2+} \), which has a 2+ charge. Because the Cu atom lost two electrons, the charge is more positive. This means that Cu was oxidized in the reaction.

\[ \text{Cu}^0(s) \rightarrow \text{Cu}^{2+}(s) + 2e^- \quad \text{Oxidation: loss of electrons by Cu} \]

At the same time, the element O in the reactants has a charge of 0, but in the CuO product, it is present as \( \text{O}^{2-} \), which has a 2− charge. Because the O atom has gained two electrons, the charge is more negative. This means that O was reduced in the reaction.

\[ \text{O}_2^0(g) + 4e^- \rightarrow 2\text{O}^{2-}(s) \quad \text{Reduction: gain of electrons by O} \]

Thus, the overall equation for the formation of CuO involves an oxidation and a reduction that occur simultaneously. In every oxidation and reduction, the number of electrons lost must be equal to the number of electrons gained. Therefore, we multiply the oxidation reaction of Cu by 2. Canceling the \( 4e^- \) on each side, we obtain the overall oxidation–reduction equation for the formation of CuO.

\[
\begin{align*}
2\text{Cu}(s) & \rightarrow 2\text{Cu}^{2+}(s) + 4e^- & \text{Oxidation} \\
\text{O}_2(g) + 4e^- & \rightarrow 2\text{O}^{2-}(s) & \text{Reduction} \\
2\text{Cu}(s) + \text{O}_2(g) & \rightarrow 2\text{CuO}(s) & \text{Oxidation–reduction equation}
\end{align*}
\]

As we see in the next reaction between zinc and copper(II) sulfate, there is always an oxidation with every reduction (see **FIGURE 8.6**).

\[ \text{Zn}(s) + \text{CuSO}_4(aq) \rightarrow \text{ZnSO}_4(aq) + \text{Cu}(s) \]

We rewrite the equation to show the atoms and ions.

\[ \text{Zn}(s) + \text{Cu}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{SO}_4^{2-}(aq) + \text{Cu}(s) \]

In this reaction, Zn atoms lose two electrons to form \( \text{Zn}^{2+} \). The increase in positive charge indicates that Zn is oxidized. At the same time, \( \text{Cu}^{2+} \) gains two electrons. The decrease in charge indicates that Cu is reduced. The \( \text{SO}_4^{2-} \) ions are **spectator ions**, which are present in both the reactants and products and do not change.

\[
\begin{align*}
\text{Zn}(s) & \rightarrow \text{Zn}^{2+}(aq) + 2e^- & \text{Oxidation of Zn} \\
\text{Cu}^{2+}(aq) + 2e^- & \rightarrow \text{Cu}(s) & \text{Reduction of Cu}^{2+}
\end{align*}
\]

In this single replacement reaction, zinc was oxidized and copper(II) ion was reduced.

**Oxidation and Reduction in Biological Systems**

Oxidation may also involve the addition of oxygen or the loss of hydrogen, and reduction may involve the loss of oxygen or the gain of hydrogen. In the cells of the body, oxidation of organic (carbon) compounds involves the transfer of hydrogen atoms (H), which are composed of
electrons and protons. For example, the oxidation of a typical biochemical molecule can involve the transfer of two hydrogen atoms (or $2H^+$ and $2e^-$) to a hydrogen ion acceptor such as the coenzyme FAD (flavin adenine dinucleotide). The coenzyme is reduced to $\text{FADH}_2$.

<table>
<thead>
<tr>
<th>Oxidation (loss of 2H)</th>
</tr>
</thead>
<tbody>
<tr>
<td>2H in biological molecule</td>
</tr>
<tr>
<td>Coenzyme FAD</td>
</tr>
<tr>
<td>Oxidized biological molecule</td>
</tr>
<tr>
<td>Coenzyme FADH$_2$</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reduction (gain of 2H)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zn$^0$ $\rightarrow$ Zn$^{2+}$</td>
</tr>
<tr>
<td>Cu$^{2+}$ $\rightarrow$ Cu$^0$ $+ 2e^-$</td>
</tr>
</tbody>
</table>

In many biochemical oxidation–reduction reactions, the transfer of hydrogen atoms is necessary for the production of energy in the cells. For example, methyl alcohol (CH$_4$O), a poisonous substance, is metabolized in the body by the following reactions:

$$\text{CH}_4\text{O} \rightarrow \text{CH}_2\text{O} + 2\text{H} \quad \text{Oxidation: loss of H atoms}$$

**Methyl alcohol**

**Formaldehyde**

The formaldehyde can be oxidized further, this time by the addition of oxygen, to produce formic acid.

$$2\text{CH}_2\text{O} + \text{O}_2 \rightarrow 2\text{CH}_2\text{O}_2 \quad \text{Oxidation: addition of O atoms}$$

**Formaldehyde**

**Formic acid**

Finally, formic acid is oxidized to carbon dioxide and water.

$$2\text{CH}_2\text{O}_2 + \text{O}_2 \rightarrow 2\text{CO}_2 + 2\text{H}_2\text{O} \quad \text{Oxidation: addition of O atoms}$$

**Formic acid**

**Carbon dioxide**

**Water**

The intermediate products of the oxidation of methyl alcohol are quite toxic, causing blindness and possibly death as they interfere with key reactions in the cells of the body.

In summary, we find that the particular definition of oxidation and reduction we use depends on the process that occurs in the reaction. All these definitions are summarized in **TABLE 8.4.** Oxidation always involves a loss of electrons, but it may also be seen as an addition of oxygen or the loss of hydrogen atoms. A reduction always involves a gain of electrons and may also be seen as the loss of oxygen or the gain of hydrogen.

**TABLE 8.4 Characteristics of Oxidation and Reduction**

<table>
<thead>
<tr>
<th>Oxidation</th>
<th>May Involve</th>
</tr>
</thead>
<tbody>
<tr>
<td>Loss of electrons</td>
<td>Addition of oxygen</td>
</tr>
<tr>
<td>Loss of hydrogen</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reduction</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Gain of electrons</td>
<td>Loss of oxygen</td>
</tr>
<tr>
<td>Gain of hydrogen</td>
<td></td>
</tr>
</tbody>
</table>
8.25 Identify each of the following as an oxidation or a reduction:
   a. \( \text{Na}^+ (aq) + e^- \rightarrow \text{Na}(s) \)
   b. \( \text{Ni}(s) \rightarrow \text{Ni}^{2+} (aq) + 2e^- \)
   c. \( \text{Cr}^{3+} (aq) + 3e^- \rightarrow \text{Cr}(s) \)
   d. \( 2\text{H}^+ (aq) + 2e^- \rightarrow \text{H}_2(g) \)

8.26 Identify each of the following as an oxidation or a reduction:
   a. \( \text{O}_2(g) + 4e^- \rightarrow 2\text{O}^{2-}(aq) \)
   b. \( \text{Ag}(s) \rightarrow \text{Ag}^+(aq) + e^- \)
   c. \( \text{Fe}^{3+} (aq) + e^- \rightarrow \text{Fe}^{2+} (aq) \)
   d. \( 2\text{Br}^-(aq) \rightarrow \text{Br}_2(l) + 2e^- \)

8.27 In each of the following reactions, identify the reactant that is oxidized and the reactant that is reduced:
   a. \( \text{Zn}(s) + \text{Cl}_2(g) \rightarrow \text{ZnCl}_2(s) \)
   b. \( \text{Cl}_2(g) + 2\text{NaBr}(aq) \rightarrow 2\text{NaCl}(aq) + \text{Br}_2(l) \)
   c. \( 2\text{PbO}(s) \rightarrow 2\text{Pb}(s) + \text{O}_2(g) \)
   d. \( 2\text{Fe}^{3+}(aq) + 2\text{Sn}^{2+}(aq) \rightarrow 2\text{Fe}^{2+}(aq) + 2\text{Sn}^{4+}(aq) \)

8.28 In each of the following reactions, identify the reactant that is oxidized and the reactant that is reduced:
   a. \( 2\text{Li}(s) + \text{F}_2(g) \rightarrow 2\text{LiF}(s) \)
   b. \( \text{Cl}_2(g) + 2\text{Kl}(aq) \rightarrow 2\text{I}_2(s) + 2\text{KCl}(aq) \)
   c. \( 2\text{Al}(s) + 3\text{Sn}^{2+}(aq) \rightarrow 2\text{Al}^{3+}(aq) + 3\text{Sn}(s) \)
   d. \( \text{Fe}(s) + \text{CuSO}_4(aq) \rightarrow \text{Cu}(s) + \text{FeSO}_4(aq) \)

Applications

8.29 In the mitochondria of human cells, energy is provided by the oxidation and reduction reactions of the iron ions in the cytochromes in electron transport. Identify each of the following reactions as an oxidation or a reduction:
   a. \( \text{Fe}^{3+} + e^- \rightarrow \text{Fe}^{2+} \)
   b. \( \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^- \)

8.30 Chlorine (Cl\(_2\)) is a strong germicide used to disinfect drinking water and to kill microbes in swimming pools. If the product is Cl\(^-\), was the elemental chlorine oxidized or reduced?

8.31 When linoleic acid, an unsaturated fatty acid, reacts with hydrogen, it forms a saturated fatty acid. Is linoleic acid oxidized or reduced in the hydrogenation reaction?

\[
\text{C}_{18}\text{H}_2\text{O}_2 + 2\text{H}_2 \rightarrow \text{C}_{18}\text{H}_{30}\text{O}_2
\]

Linoleic acid

8.32 In one of the reactions in the citric acid cycle, which provides energy, succinic acid is converted to fumaric acid.

\[
\text{C}_4\text{H}_6\text{O}_4 \rightarrow \text{C}_4\text{H}_4\text{O}_4 + 2\text{H}
\]

Succinic acid  Fumaric acid

The reaction is accompanied by the reaction of a coenzyme, flavin adenine dinucleotide (FAD).

\[
\text{FAD} + 2\text{H} \rightarrow \text{FADH}_2
\]

a. Is succinic acid oxidized or reduced?
   b. Is FAD oxidized or reduced?
   c. Why would the two reactions occur together?

Follow Up

IMPROVING NATALIE’S OVERALL FITNESS

Natalie’s test results indicate that she has a blood oxygen level of 89%. The normal values for pulse oximeter readings are 95% to 100%, which means that Natalie’s \( O_2 \) saturation is low. Thus, Natalie may be hypoxic. This may be the reason she has noticed a shortness of breath and a dry cough. Her doctor diagnosed her with interstitial lung disease, which is scarring of the tissue of the lungs.

Angela teaches Natalie to inhale and exhale slower and deeper to fill the lungs with more air and thus more oxygen. Angela also develops a workout program with the goal of increasing Natalie’s overall fitness level. During the exercises, Angela continues to monitor Natalie’s heart rate, blood \( O_2 \) level, and blood pressure to ensure that Natalie is exercising at a level that will enable her to become stronger without breaking down muscle due to a lack of oxygen.

Applications

8.33 a. During cellular respiration, \( \text{C}_6\text{H}_12\text{O}_6 \) (glucose) in the cells undergoes combustion. Write and balance the chemical equation for reaction of glucose in the human body.

   \[
   \text{C}_6\text{H}_12\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}
   \]

b. In plants, carbon dioxide and oxygen gases are converted to glucose \( (\text{C}_6\text{H}_12\text{O}_6) \) and water. Write and balance the chemical equation for production of glucose in plants.

   \[
   6\text{CO}_2 + 6\text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_12\text{O}_6 + 6\text{O}_2
   \]

8.34 Fatty acids undergo reaction with \( O_2 \) and form \( \text{CO}_2 \) and \( \text{H}_2\text{O} \) when utilized for energy in the body.

a. Write and balance the equation for the combustion of the fatty acid capric acid, \( \text{C}_{16}\text{H}_{32}\text{O}_2 \).

   \[
   2\text{C}_{16}\text{H}_{32}\text{O}_2 + 64\text{O}_2 \rightarrow 32\text{CO}_2 + 32\text{H}_2\text{O}
   \]

b. Write and balance the equation for the combustion of the fatty acid myristic acid, \( \text{C}_{14}\text{H}_{28}\text{O}_2 \).

   \[
   \text{C}_{14}\text{H}_{28}\text{O}_2 + 21\text{O}_2 \rightarrow 14\text{CO}_2 + 14\text{H}_2\text{O}
   \]
CHEMICAL REACTIONS

show a chemical change between

Reactants and Products

are balanced with

Coefficients
to give

Equal Numbers of Atoms on Each Side

are classified as

Combination, Decomposition, Single Replacement, Double Replacement, and Combustion

and those with

Loss or Gain of Electrons

are also

Oxidation–Reduction Reactions

CHAPTER REVIEW

8.1 Equations for Chemical Reactions

LEARNING GOAL Identify a balanced chemical equation; determine the number of atoms in the reactants and products.

• A chemical reaction occurs when the atoms of the initial substances rearrange to form new substances.
• A chemical equation shows the formulas of the substances that react on the left side of a reaction arrow and the products that form on the right side of the reaction arrow.

8.2 Balancing a Chemical Equation

LEARNING GOAL Write a balanced chemical equation from the formulas of the reactants and products for a chemical reaction.

• A chemical equation is balanced by writing coefficients, small whole numbers, in front of formulas to equalize the atoms of each of the elements in the reactants and the products.

8.3 Types of Chemical Reactions

LEARNING GOAL Identify a chemical reaction as a combination, decomposition, single replacement, double replacement, or combustion.

• Many chemical reactions can be organized by reaction type: combination, decomposition, single replacement, double replacement, or combustion.

8.4 Oxidation–Reduction Reactions

LEARNING GOAL Define the terms oxidation and reduction; identify the reactants oxidized and reduced.

• When electrons are transferred in a reaction, it is an oxidation–reduction reaction.
• One reactant loses electrons, and another reactant gains electrons.
• Overall, the number of electrons lost and gained is equal.

KEY TERMS

balanced equation The final form of a chemical equation that shows the same number of atoms of each element in the reactants and products.
chemical equation A shorthand way to represent a chemical reaction using chemical formulas to indicate the reactants and products and coefficients to show reacting ratios.
coefficients Whole numbers placed in front of the formulas to balance the number of atoms or moles of atoms of each element on both sides of an equation.
combination reaction A chemical reaction in which reactants combine to form a single product.
combustion reaction A chemical reaction in which a fuel containing carbon and hydrogen reacts with oxygen to produce CO₂, H₂O, and energy.
decomposition reaction A reaction in which a single reactant splits into two or more simpler substances.
double replacement reaction A reaction in which the positive ions in the reacting compounds exchange places.
oxidation  The loss of electrons by a substance. Oxidation may involve the addition of oxygen or the loss of hydrogen.

oxidation–reduction reaction A reaction in which the oxidation of one reactant is always accompanied by the reduction of another reactant.

products The substances formed as a result of a chemical reaction.

reactants The initial substances that undergo change in a chemical reaction.

reduction The gain of electrons by a substance. Reduction may involve the loss of oxygen or the gain of hydrogen.

single replacement reaction A reaction in which one element replaces a different element in a compound.

CORE CHEMISTRY SKILLS

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

Balancing a Chemical Equation (8.2)

- In a balanced chemical equation, whole numbers called coefficients multiply each of the atoms in the chemical formulas so that the number of each type of atom in the reactants is equal to the number of the same type of atom in the products.

Example: Balance the following chemical equation:

\[ \text{SnCl}_2(s) + \text{H}_2\text{O}(l) \rightarrow \text{Sn(OH)}_2(s) + \text{HCl}(aq) \]

Answer: When we compare the atoms on the reactant side and the product side, we see that there are more Cl atoms in the reactants and more O and H atoms in the products.

To balance the equation, we need to use coefficients in front of the formulas containing the Cl atoms, H atoms, and O atoms.

- Place a 4 in front of the formula HCl to give a total of 8 H atoms and 4 Cl atoms in the products.

\[ \text{SnCl}_2(s) + \text{H}_2\text{O}(l) \rightarrow 4\text{Sn(OH)}_2(s) + 4\text{HCl}(aq) \]

- Place a 2 in front of the formula HCl to give 4 H atoms and 4 O atoms in the reactants.

\[ \text{SnCl}_2(s) + 2\text{H}_2\text{O}(l) \rightarrow 4\text{Sn(OH)}_2(s) + 4\text{HCl}(aq) \]

- The total number of Sn (1), Cl (4), H (8), and O (4) atoms is now equal on both sides of the equation.

\[ 4\text{SnCl}_2(s) + 8\text{H}_2\text{O}(l) \rightarrow 4\text{Sn(OH)}_2(s) + 8\text{HCl}(aq) \]

Classifying Types of Chemical Reactions (8.3)

- Chemical reactions are classified by identifying general patterns in their equations.

- In a combination reaction, two or more elements or compounds bond to form one product.

- In a decomposition reaction, a single reactant splits into two or more products.

- In a single replacement reaction, an uncombined element takes the place of an element in a compound.

- In a double replacement reaction, the positive ions in the reacting compounds switch places.

- In a combustion reaction, a carbon-containing compound that is the fuel burns in oxygen from the air to produce carbon dioxide (CO₂), water (H₂O), and energy.

Example: Classify the type of the following reaction:

\[ 2\text{Al(s)} + \text{Fe}_2\text{O}_3(s) \rightarrow \Delta \text{Al}_2\text{O}_3(s) + 2\text{Fe(l)} \]

Answer: The iron in iron(III) oxide is replaced by aluminum, which makes this a single replacement reaction.

Identifying Oxidized and Reduced Substances (8.4)

- In an oxidation–reduction reaction (abbreviated redox), one reactant is oxidized when it loses electrons, and another reactant is reduced when it gains electrons.

- Oxidation is the loss of electrons; reduction is the gain of electrons.

Example: For the following redox reaction, identify the reactant that is oxidized, and the reactant that is reduced:

\[ \text{Fe(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Fe}^{2+}(aq) + \text{Cu(s)} \]

Answer: \[ \text{Fe}^{0}(s) \rightarrow \text{Fe}^{2+}(aq) + 2 e^- \]

Fe loses electrons; it is oxidized.

\[ \text{Cu}^{2+}(aq) + 2 e^- \rightarrow \text{Cu}^{0}(s) \]

Cu²⁺ gains electrons; it is reduced.

UNDERSTANDING THE CONCEPTS

The chapter sections to review are shown in parentheses at the end of each question.

8.35 Balance each of the following by adding coefficients, and identify the type of reaction for each: (8.1, 8.2)

a. \[ \text{} + \text{} \rightarrow \text{} \text{} + \text{} \]

b. \[ \text{} \rightarrow \text{} \text{} + \text{} \]

8.36 Balance each of the following by adding coefficients, and identify the type of reaction for each: (8.1, 8.2)

a. \[ \text{} \text{} \rightarrow \text{} \text{} \text{} \]

b. \[ \text{} \text{} + \text{} \rightarrow \text{} \text{} \]

8.37 If red spheres represent oxygen atoms, blue spheres represent sulfur atoms, and all the molecules are gases, (8.1, 8.2)

Reactants Products

a. write the formula for each of the reactants and products.

b. write a balanced equation for the reaction.

c. indicate the type of reaction as combination, decomposition, single replacement, double replacement, or combustion.
8.38 If purple spheres represent iodine atoms, white spheres represent chlorine atoms, the reacting molecules are solid, and the products are gases, (8.1, 8.2)

Reactants  
[Diagram of purple and white spheres]  
Products  
[Diagram of purple and white spheres]

a. write the formula for each of the reactants and products.  
b. write a balanced equation for the reaction.  
c. indicate the type of reaction as combination, decomposition, single replacement, double replacement, or combustion.

8.39 If blue spheres represent nitrogen atoms, purple spheres represent bromine atoms, and all the molecules are gases, (8.1, 8.2)

Reactants  
[Diagram of blue and purple spheres]  
Products  
[Diagram of blue and purple spheres]

a. write the formula for each of the reactants and products.  
b. write a balanced equation for the reaction.  
c. indicate the type of reaction as combination, decomposition, single replacement, double replacement, or combustion.

8.40 If green spheres represent chlorine atoms, yellow-green spheres represent fluorine atoms, white spheres represent hydrogen atoms, and all the molecules are gases, (8.1, 8.2, 8.3)

Reactants  
[Diagram of green, yellow-green, and white spheres]  
Products  
[Diagram of green, yellow-green, and white spheres]

a. write the formula for each of the reactants and products.  
b. write a balanced equation for the reaction.  
c. indicate the type of reaction as combination, decomposition, single replacement, double replacement, or combustion.

8.41 If green spheres represent chlorine atoms, red spheres represent oxygen atoms, and all the molecules are gases, (8.1, 8.2, 8.3)

Reactants  
[Diagram of green and red spheres]  
Products  
[Diagram of green and red spheres]

a. write the formula for each of the reactants and products.  
b. write a balanced equation for the reaction.  
c. indicate the type of reaction as combination, decomposition, single replacement, double replacement, or combustion.

8.42 If blue spheres represent nitrogen atoms, purple spheres represent iodine atoms, the reacting molecules are gases, and the products are solid, (8.1, 8.2, 8.3)

Reactants  
[Diagram of blue and purple spheres]  
Products  
[Diagram of blue and purple spheres]

a. write the formula for each of the reactants and products.  
b. write a balanced equation for the reaction.

8.43 Identify the type of reaction for each of the following as combination, decomposition, single replacement, double replacement, or combustion: (8.3)

a. Potassium reacts with chlorine gas to form potassium chloride.  
b. Potassium metal and chlorine gas are obtained from electrolysis of molten potassium chloride.  
c. Heating starch in oxygen produces carbon dioxide and water.

8.44 Identify the type of reaction for each of the following as combination, decomposition, single replacement, double replacement, or combustion: (8.3)

a. A compound breaks apart into its elements.  
b. Copper and bromine form copper(II) bromide.  
c. Iron(II) sulfite breaks down to iron(II) oxide and sulfur dioxide.  
d. Silver ion from AgNO₃(aq) forms a solid with bromide ion from KBr(aq).
8.45 Balance each of the following chemical equations, and identify the type of reaction: (8.1, 8.2, 8.3)
   a. \( \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \)
   b. \( \text{C}_2\text{H}_6\text{O}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) \)
   c. \( \text{Sb}(s) + \text{Cl}_2(g) \rightarrow \text{SbCl}_3(s) \)
   d. \( \text{Ni}_3(s) \rightarrow 2\text{Ni}(g) + \text{I}_2(g) \)
   e. \( \text{KBr}(aq) + \text{Cl}_2(aq) \rightarrow \text{KCl}(aq) + \text{Br}_2(l) \)
   f. \( \text{Fe}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{H}_2(g) + \text{Fe}_2\text{SO}_3(aq) \)
   g. \( \text{Al}_2\text{SO}_4(aq) + \text{NaOH}(aq) \rightarrow \text{Al(OH)}_3(s) + \text{Na}_2\text{SO}_4(aq) \)

8.46 Balance each of the following chemical equations, and identify the type of reaction: (8.1, 8.2, 8.3)
   a. \( \text{Si}_3\text{N}_4(s) \rightarrow \text{Si}(s) + \text{N}_2(g) \)
   b. \( \text{Mg}(s) + \text{N}_2(g) \rightarrow \text{Mg}_3\text{N}_2(s) \)
   c. \( \text{Al}(s) + \text{H}_2\text{PO}_4(aq) \rightarrow \text{H}_2(g) + \text{AlPO}_4(aq) \)
   d. \( \text{C}_2\text{H}_6(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) \)
   e. \( \text{Cr}_2\text{O}_3(s) + \text{H}_2(g) \rightarrow \text{Cr}(s) + \text{H}_2\text{O}(g) \)
   f. \( \text{Al}(s) + \text{Cl}_2(g) \rightarrow \text{AlCl}_3(s) \)
   g. \( \text{MgCl}_2(aq) + \text{AgNO}_3(aq) \rightarrow \text{AgCl}(s) + \text{Mg(NO}_3)_2(aq) \)

8.47 Predict the products and write a balanced equation for each of the following: (8.1, 8.2, 8.3)
   a. single replacement:
      \( \text{Zn}(s) + \text{HCl}(aq) \rightarrow _____ + _____ \)
   b. decomposition:
      \( \text{BaCO}_3(s) \rightarrow _____ + _____ \)
   c. double replacement:
      \( \text{NaOH}(aq) + \text{HCl}(aq) \rightarrow _____ + _____ \)
   d. combination:
      \( \text{Al}(s) + \text{F}_2(g) \rightarrow _____ \)

8.48 Predict the products and write an equation for each of the following: (8.1, 8.2, 8.3)
   a. electrification:
      \( \text{NaCl}(s) \rightarrow _____ + _____ \)
   b. combination:
      \( \text{Ca}(s) + \text{Br}_2(l) \rightarrow _____ \)

8.49 Write a balanced equation for each of the following reactions and identify the type of reaction: (8.1, 8.2, 8.3)
   a. \( \text{Na}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \)
   b. \( \text{Mg}_2(g) + 3\text{Mg}(g) \rightarrow \text{Mg}_2\text{N}_2(g) \)
   c. \( 2\text{Al}(s) + 6\text{HCl}(aq) \rightarrow 2\text{AlCl}_3(aq) + 3\text{H}_2(g) \)

8.50 Write a balanced equation for each of the following reactions and identify the type of reaction: (8.1, 8.2, 8.3)
   a. \( \text{Na}_2\text{O}(s) + \text{H}_2\text{O}(l) \rightarrow 2\text{NaOH}(aq) + \text{H}_2(g) \)
   b. \( 4\text{FeS}(s) + 7\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) + 4\text{SO}_2(g) \)
   c. \( \text{Zn}(s) + \text{Pb(NO}_3)_2(aq) \rightarrow \text{Zn(NO}_3)_2(aq) + \text{Pb}(s) \)

**CHALLENGE QUESTIONS**

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

8.53 Balance each of the following chemical equations, and identify the type of reaction: (8.1, 8.2, 8.3)
   a. \( \text{MgCO}_3(s) \rightarrow \text{MgO}(s) + \text{CO}_2(g) \)
   b. \( \text{C}_2\text{H}_6\text{O}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) \)
   c. \( \text{Al}(s) + \text{CuCl}_2(aq) \rightarrow \text{AlCl}_3(aq) + \text{Cu}(s) \)
   d. \( \text{AgNO}_3(aq) + \text{MgCl}_2(aq) \rightarrow \text{AgCl}(s) + \text{Mg(NO}_3)_2(aq) \)

8.54 Balance each of the following chemical equations, and identify the type of reaction: (8.1, 8.2, 8.3)
   a. \( \text{Ba(NO}_3)_2(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + \text{HNO}_3(aq) \)
   b. \( \text{KCN}(aq) + \text{H}_2\text{PO}_4(aq) \rightarrow \text{HCN}(aq) + \text{K}_2\text{PO}_4(aq) \)
   c. \( \text{PbCl}_2(aq) + \text{Al}(s) \rightarrow \text{AlCl}_3(aq) + \text{Pb}(s) \)
   d. \( \text{C}_2\text{H}_2(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) \)

8.55 Complete and balance each of the following chemical equations: (8.1, 8.2, 8.3)
   a. Single replacement:
      \( \text{Fe}_3\text{O}_4(s) + \text{H}_2(g) \rightarrow _____ \)
   b. Combustion:
      \( \text{C}_2\text{H}_2\text{O}(g) + \text{O}_2(g) \rightarrow _____ \)
   c. Combination:
      \( \text{Al}(s) + \text{O}_2(g) \rightarrow _____ \)
   d. Double replacement:
      \( \text{NaOH}(aq) + \text{ZnSO}_4(aq) \rightarrow _____ \)

8.56 Complete and balance each of the following chemical equations: (8.1, 8.2, 8.3)
   a. Decomposition:
      \( \text{H}_2\text{O}(s) \rightarrow _____ \)
   b. Double replacement:
      \( \text{BaCl}_2(aq) + \text{AgNO}_3(aq) \rightarrow _____ \)
   c. Single replacement:
      \( \text{Ca}(s) + \text{AlCl}_3(s) \rightarrow _____ \)
   d. Combination:
      \( \text{Mg}(s) + \text{N}_2(g) \rightarrow _____ \)

8.57 Write the correct formulas for the reactants and products, the balanced equation for each of the following reaction descriptions, and identify each type of reaction: (8.1, 8.2, 8.3)
   a. An aqueous solution of lead(II) nitrate is mixed with aqueous sodium phosphate to produce solid lead(II) phosphate and aqueous sodium nitrate.
   b. Gallium metal heated in oxygen gas forms solid gallium(III) oxide.
   c. When solid sodium nitrate is heated, solid sodium nitrate and oxygen gas are produced.

8.58 Write the correct formulas for the reactants and products, the balanced equation for each of the following reaction descriptions, and identify each type of reaction: (8.1, 8.2, 8.3)
   a. Solid bismuth(III) oxide and solid carbon react to form bismuth metal and carbon monoxide gas.
   b. Solid sodium bicarbonate is heated and forms solid sodium carbonate, gaseous carbon dioxide, and liquid water.
   c. Liquid hexane, \( \text{C}_6\text{H}_{14} \), reacts with oxygen gas to form two gaseous products: carbon dioxide and liquid water.
8.59 In the following diagram, if blue spheres are the element X and yellow spheres are the element Y: (8.1, 8.2, 8.3)

**Reactants**

**Products**

- a. Write the formula for each of the reactants and products.
- b. Write a balanced equation for the reaction.
- c. Indicate the type of reaction as combination, decomposition, single replacement, double replacement, or combustion.

**ANSWERS**

Answers to Selected Questions and Problems

8.1 a. reactants/products: 7 O atoms
   b. reactants/products: 4 O atoms
   c. reactants/products: 10 O atoms
   d. reactants/products: 4 O atoms

8.3 a. not balanced  
   b. not balanced  
   c. balanced  
   d. balanced

8.5 a. 2 Na atoms, 2 Cl atoms
   b. 1 P atom, 3 Cl atoms, 6 H atoms
   c. 4 P atoms, 16 O atoms, 12 H atoms

8.7 a. \( \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \)
   b. \( 2\text{HgO}(s) \rightarrow 2\text{Hg}(l) + \text{O}_2(g) \)
   c. \( 4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \)
   d. \( 2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s) \)
   e. \( 2\text{Cu}_2\text{O}(s) + \text{O}_2(g) \rightarrow 4\text{CuO}(s) \)

8.9 a. \( \text{Mg}(s) + 2\text{AgNO}_3(aq) \rightarrow 2\text{Ag}(s) + \text{Mg(NO}_3)_2(aq) \)
   b. \( \text{CuCO}_3(s) \rightarrow \text{CuO}(s) + \text{CO}_2(g) \)
   c. \( 2\text{Al}(s) + 3\text{CuSO}_4(aq) \rightarrow 3\text{Cu}(s) + \text{Al}_2\text{(SO}_4)_3(aq) \)
   d. \( \text{Pb(NO}_3)_2(aq) + 2\text{NaCl}(aq) \rightarrow \text{PbCl}_2(s) + 2\text{NaNO}_3(aq) \)
   e. \( 2\text{Al}(s) + 6\text{HCl}(aq) \rightarrow 3\text{H}_2(g) + 2\text{AlCl}_3(aq) \)

8.11 a. \( \text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g) \)
   b. \( 2\text{Li}_2\text{N}(s) \rightarrow 6\text{Li}(s) + \text{N}_2(g) \)
   c. \( 2\text{Al}(s) + 6\text{HBr}(aq) \rightarrow 3\text{H}_2(g) + 2\text{AlBr}_3(aq) \)
   d. \( 3\text{Ba}(\text{OH})_2(aq) + 2\text{Na}_2\text{PO}_4(aq) \rightarrow \text{Ba}_3\text{(PO}_4)_2(s) + 6\text{NaOH}(aq) \)
   e. \( \text{As}_2\text{S}_6(s) + 9\text{O}_2(g) \rightarrow 2\text{As}_2\text{O}_3(s) + 6\text{SO}_2(g) \)

8.13 a. \( 2\text{Li}(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{H}_2(g) + 2\text{LiOH}(aq) \)
   b. \( 2\text{P}(s) + 5\text{Cl}_2(g) \rightarrow 2\text{PCl}_5(s) \)
   c. \( \text{FeO}(s) + \text{CO}(g) \rightarrow \text{Fe}(s) + \text{CO}_2(g) \)
   d. \( 2\text{C}_2\text{H}_2(l) + 15\text{O}_2(g) \rightarrow 10\text{CO}_2(g) + 10\text{H}_2\text{O}(g) \)
   e. \( 3\text{H}_2\text{S}(g) + 2\text{FeCl}_3(s) \rightarrow \text{Fe}_2\text{S}_3(s) + 6\text{HCl}(g) \)

8.15 \( \text{NH}_2\text{NO}_3(s) \rightarrow 2\text{H}_2\text{O}(g) + \text{N}_2(g) \)

8.17 \( 2\text{C}_2\text{H}_2\text{NO}_2(aq) + 6\text{O}_2(g) \rightarrow 5\text{CO}_2(g) + 5\text{H}_2\text{O}(l) + \text{CH}_3\text{N}_2\text{O}(aq) \)

8.19 a. decomposition  
   b. single replacement  
   c. combustion  
   d. double replacement  
   e. combination

8.21 a. combination  
   b. single replacement  
   c. decomposition  
   d. double replacement  
   e. combustion

8.23 a. \( \text{Mg}(s) + \text{Cl}_2(g) \rightarrow \text{MgCl}_2(s) \)
   b. \( 2\text{HBr}(g) \rightarrow \text{H}_2(g) + \text{Br}_2(l) \)
   c. \( \text{Mg}(s) + \text{Zn(NO}_3)_2(aq) \rightarrow \text{Zn}(s) + \text{Mg(NO}_3)_2(aq) \)
   d. \( \text{K}_2\text{S}(aq) + \text{Pb(NO}_3)_2(aq) \rightarrow \text{PbS}(s) + 2\text{KNO}_3(aq) \)
   e. \( 2\text{C}_2\text{H}_2(g) + 7\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 6\text{H}_2\text{O}(g) \)

8.25 a. reduction  
   b. oxidation  
   c. reduction  
   d. reduction

8.27 a. Zn is oxidized; Cl₂ is reduced.
   b. The Br⁻ in NaBr is oxidized; Cl₂ is reduced.
   c. The O²⁻ in PbO is oxidized; the Pb³⁺ in PbO is reduced.
   d. Sn²⁺ is oxidized; Fe³⁺ is reduced.

8.29 a. reduction  
   b. oxidation

8.31 Linoeic acid gains hydrogen atoms and is reduced.

8.33 a. \( \text{C}_2\text{H}_4\text{O}_2(aq) + 6\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 6\text{H}_2\text{O}(l) \)
   b. \( 6\text{CO}_2(g) + 6\text{H}_2\text{O}(l) \rightarrow \text{C}_2\text{H}_4\text{O}_2(aq) + 6\text{O}_2(g) \)

8.35 a. 1, 2, 1 single replacement
   b. 1, 1, 1 decomposition

8.37 a. reactants SO and O₂; product SO
   b. 2SO(g) + O₂(g) \rightarrow 2SO₂(g)
   c. combination

8.39 a. reactant NBr₃; product N₂ and H₂
   b. 2NBr₃(g) \rightarrow N₂(g) + 3Br₂(g)
   c. decomposition

8.41 a. reactants Cl₂ and O₂; product OCl₂
   b. 2Cl₂(g) + O₂(g) \rightarrow 2OCl₂(g)
   c. combination

8.43 a. combination
   b. decomposition
   c. combustion

8.45 a. \( \text{NH}_4(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \)
   b. \( \text{C}_2\text{H}_4\text{O}(g) + \text{H}_2\text{O}(l) \rightarrow 4\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \)
   c. 2Sb(s) + 3Cl₂(g) \rightarrow 2SbCl₅(s)
   d. 2Ni(s) \rightarrow Ni₂(g) + 3I₂(g) decomposition
   e. \( 2\text{KBr}(aq) + \text{Cl}_2(aq) \rightarrow 2\text{KCl}(aq) + \text{Br}_2(l) \)
   single replacement
f. $2\text{Fe(s)} + 3\text{H}_2\text{SO}_4(aq) \rightarrow 3\text{H}_2(g) + \text{Fe}_2(\text{SO}_4)_3(aq)$
   \hspace{1cm} \text{single replacement}

\text{g. } \text{Al}_2(\text{SO}_4)_3(aq) + 6\text{NaOH}(aq) \rightarrow 2\text{Al(OH)}_3(s) + 3\text{Na}_2\text{SO}_4(aq)$ \text{double replacement}

8.47 a. $\text{Zn(s)} + 2\text{HCl}(aq) \rightarrow \text{H}_2(g) + \text{ZnCl}_2(aq)$
b. $\text{BaCO}_3(s) \xrightarrow{\Delta} \text{BaO(s)} + \text{CO}_2(g)$
c. $\text{NaOH}(aq) + \text{HCl}(aq) \rightarrow \text{H}_2\text{O(l)} + \text{NaCl}(aq)$
d. $2\text{Al(s)} + 3\text{F}_2(g) \rightarrow 2\text{AlF}_3(s)$

8.49 a. $4\text{Na(s)} + \text{O}_2(g) \rightarrow 2\text{Na}_2(s)$ \text{combination}
b. $\text{NaCl}(aq) + \text{AgNO}_3(aq) \rightarrow \text{AgCl(s)} + \text{NaNO}_3(aq)$ \text{double replacement}
c. $\text{C}_2\text{H}_5\text{O}(l) + 3\text{O}_2(g) \xrightarrow{\Delta} 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g)$ \text{combustion}

8.51 a. $\text{N}_2$ is oxidized; $\text{H}_2$ is reduced.
b. $\text{N}_2$ is oxidized; $\text{Mg}$ is reduced.
c. $\text{Al}$ is oxidized; $\text{H}$ is reduced.

8.53 a. already balanced, \text{decomposition}
b. $\text{C}_2\text{H}_2\text{O}_6 + 6\text{O}_2 \rightarrow 2\text{CO}_2 + 6\text{H}_2\text{O}$ \text{combustion}
c. $2\text{Al} + 3\text{CuCl}_2 \rightarrow 2\text{AlCl}_3 + 3\text{Cu}$ \text{single displacement}
d. $2\text{AgNO}_3(aq) + \text{MgCl}_2(aq) \rightarrow 2\text{AgCl(s)} + \text{Mg(NO}_3)_2(aq)$ \text{double displacement}

8.55 a. $\text{Fe}_3\text{O}_4(s) + 4\text{H}_2(g) \rightarrow 3\text{Fe(s)} + 4\text{H}_2\text{O}(g)$
b. $2\text{C}_2\text{H}_2\text{O}(g) + 13\text{O}_2(g) \xrightarrow{\Delta} 8\text{CO}_2(g) + 10\text{H}_2\text{O}(g)$
c. $4\text{Al(s)} + 3\text{O}_2(g) \rightarrow 2\text{Al}_2\text{O}_3(s)$
d. $\text{NaOH}(aq) + \text{ZnSO}_4(aq) \rightarrow \text{Zn(OH)}_2(s) + \text{Na}_2\text{SO}_4(aq)$

8.57 a. $3\text{Ph(NO}_3)_2(aq) + 2\text{Na}_3\text{PO}_4(aq) \rightarrow \text{Pb}_3(\text{PO}_4)_2(s) + 6\text{NaNO}_3(aq)$ \text{double replacement}
b. $4\text{Ga(s)} + 3\text{O}_2(g) \xrightarrow{\Delta} 2\text{Ga}_2\text{O}_3(s)$ \text{combination}
c. $2\text{NaNO}_3(s) \xrightarrow{\Delta} 2\text{NaNO}_2(s) + \text{O}_2(g)$ \text{decomposition}

8.59 a. reactants: $X$ and $Y_2$; product: $XY_3$
b. $2X + 3Y_2 \rightarrow 2XY_3$
c. combination
LANCE, AN ENVIRONMENTAL scientist, is collecting soil and water samples at a nearby farm to test for the presence and concentration of any pesticides and pharmaceuticals. Farmers use pesticides to increase food production and pharmaceuticals to treat and prevent animal-related diseases. Due to the common use of these chemicals, they may pass into the soil and water supply, potentially contaminating the environment and causing health problems.

Recently, a farmer treated his cotton and bean fields with a pesticide called Sevin (carbaryl). A few days later, Lance collected samples of soil and water and detected small amounts of Sevin. In humans, Sevin, which is an acetylcholinesterase inhibitor, can cause headaches, nausea, and paralysis of the respiratory system. Because it is very soluble, it is important that the residue pesticide does not contaminate the water supply. Lance advised the farmer to decrease the amount of pesticide he uses on his crops to reduce the pesticide level. He indicates that he will return in a week to retest the soil and water for Sevin.

**CAREER**

**Environmental Scientist**

Environmental science is a multidisciplinary field that combines chemistry, biology, ecology, and geology to study environmental problems. Environmental scientists monitor environmental pollution to protect the health of the public. By using specialized equipment, environmental scientists measure pollution levels in soil, air, and water, as well as noise and radiation levels. They can specialize in a specific area, such as air quality or hazardous and solid waste. For instance, air-quality experts monitor indoor air for allergens, mold, and toxins; they measure outdoor air pollutants created by businesses, vehicles, and agriculture. Since environmental scientists obtain samples containing potentially hazardous materials, they must be knowledgeable about safety protocols and wear personal protective equipment. They may also recommend methods to diminish various pollutants, and may assist in cleanup and remediation efforts.
9.1 Conservation of Mass

**LEARNING GOAL** Calculate the total mass of reactants and the total mass of products in a balanced chemical equation.

In any chemical reaction, the total amount of matter in the reactants is equal to the total amount of matter in the products. Thus, the total mass of all the reactants must be equal to the total mass of all the products. This is known as the **law of conservation of mass**, which states that there is no change in the total mass of the substances reacting in a balanced chemical reaction. Thus, no material is lost or gained as original substances are changed to new substances.

For example, tarnish (Ag₂S) forms when silver reacts with sulfur to form silver sulfide.

$$2\text{Ag}(s) + \text{S}(s) \rightarrow \text{Ag}_2\text{S}(s)$$

In this reaction, the number of silver atoms that react is twice the number of sulfur atoms. When 200 silver atoms react, 100 sulfur atoms are required. However, in the actual...
chemical reaction, many more atoms of both silver and sulfur would react. If we are dealing with molar amounts, then the coefficients in the equation can be interpreted in terms of moles. Thus, 2 mol of silver reacts with 1 mol of sulfur to produce 1 mol of Ag₂S. Because the molar mass of each can be determined, the moles of Ag, S, and Ag₂S can also be stated in terms of mass in grams of each. Thus, 215.8 g of Ag and 32.1 g of S react to form 247.9 g of Ag₂S. The total mass of the reactants (247.9 g) is equal to the mass of the product (247.9 g). The various ways in which a chemical equation can be interpreted are seen in **TABLE 9.1**.

**TABLE 9.1** Information Available from a Balanced Equation

<table>
<thead>
<tr>
<th></th>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Equation</strong></td>
<td>2Ag(s) + S(s) → Ag₂S(s)</td>
<td></td>
</tr>
<tr>
<td><strong>Atoms/Formula Units</strong></td>
<td>2 Ag atoms + 1 S atom → 1 Ag₂S formula unit</td>
<td></td>
</tr>
<tr>
<td><strong>Avogadro’s Number of Atoms</strong></td>
<td>2(6.022 × 10²³) Ag atoms + 1(6.022 × 10²³) S atoms → 1(6.022 × 10²³) Ag₂S formula units</td>
<td></td>
</tr>
<tr>
<td><strong>Moles</strong></td>
<td>2 mol of Ag + 1 mol of S → 1 mol of Ag₂S</td>
<td></td>
</tr>
<tr>
<td><strong>Mass (g)</strong></td>
<td>2(107.9 g) of Ag + 1(32.07 g) of S → 1(247.9 g) of Ag₂S</td>
<td></td>
</tr>
<tr>
<td><strong>Total Mass (g)</strong></td>
<td>247.9 g</td>
<td>247.9 g</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 9.1 Conservation of Mass**

The combustion of methane (CH₄) with oxygen produces carbon dioxide, water, and energy. Calculate the total mass of the reactants and the products for the following equation when 1 mol of CH₄ reacts:

\[ CH_4(g) + 2O_2(g) \xrightarrow[\Delta]{} CO_2(g) + 2H_2O(g) \]

**TRY IT FIRST**

**SOLUTION**

Interpreting the coefficients in the equation as the number of moles of each substance and multiplying by its molar mass gives the total mass of reactants and products. The quantities of moles are exact because the coefficients in the balanced equation are exact.

<table>
<thead>
<tr>
<th></th>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Equation</strong></td>
<td>CH₄(g) + 2O₂(g) → CO₂(g) + 2H₂O(g)</td>
<td></td>
</tr>
<tr>
<td><strong>Moles</strong></td>
<td>1 mol of CH₄ + 2 mol of O₂ → 1 mol of CO₂ + 2 mol of H₂O</td>
<td></td>
</tr>
<tr>
<td><strong>Mass</strong></td>
<td>16.04 g of CH₄ + 64.00 g of O₂ → 44.01 g of CO₂ + 36.03 g of H₂O</td>
<td></td>
</tr>
<tr>
<td><strong>Total Mass</strong></td>
<td>80.04 g of reactants = 80.04 g of products</td>
<td></td>
</tr>
</tbody>
</table>

**STUDY CHECK 9.1**

Calculate the total mass of the reactants and the products for the following equation:

\[ 4K(s) + O_2(g) \rightarrow 2K_2O(s) \]

**ANSWER**

188.4 g of reactants and 188.4 g of products
Questions and Problems

9.1 Conservation of Mass

LEARNING GOAL  Calculate the total mass of reactants and the total mass of products in a balanced chemical equation.

9.1 Calculate the total mass of the reactants and the products for each of the following equations:

- a. \(2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g)\)
- b. \(4\text{P}(s) + 5\text{O}_2(g) \rightarrow 2\text{P}_2\text{O}_5(s)\)

9.2 Calculate the total mass of the reactants and the products for each of the following equations:

- a. \(2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s)\)
- b. \(4\text{HCl}(g) + \text{O}_2(g) \rightarrow 2\text{Cl}_2(g) + 2\text{H}_2\text{O}(g)\)

9.2 Calculating Moles Using Mole–Mole Factors

LEARNING GOAL  Use a mole–mole factor from a balanced chemical equation to calculate the number of moles of another substance in the reaction.

When iron reacts with sulfur, the product is iron(III) sulfide.

\[2\text{Fe}(s) + 3\text{S}(s) \rightarrow \text{Fe}_2\text{S}_3(s)\]

From the balanced equation, we see that 2 mol of iron reacts with 3 mol of sulfur to form 1 mol of iron(III) sulfide. Actually, any amount of iron or sulfur may be used, but the ratio of iron reacting with sulfur will always be the same. From the coefficients, we can write mole–mole factors between reactants and between reactants and products. The coefficients used in the mole–mole factors are exact numbers; they do not limit the number of significant figures.

- \(\frac{2 \text{ mol Fe}}{3 \text{ mol S}}\) and \(\frac{3 \text{ mol S}}{2 \text{ mol Fe}}\)
- \(\frac{2 \text{ mol Fe}}{1 \text{ mol } \text{Fe}_2\text{S}_3}\) and \(\frac{1 \text{ mol } \text{Fe}_2\text{S}_3}{2 \text{ mol Fe}}\)
- \(\frac{3 \text{ mol S}}{1 \text{ mol } \text{Fe}_2\text{S}_3}\) and \(\frac{1 \text{ mol } \text{Fe}_2\text{S}_3}{3 \text{ mol S}}\)

Using Mole–Mole Factors in Calculations

Whenever you prepare a recipe, adjust an engine for the proper mixture of fuel and air, or prepare medicines in a pharmaceutical laboratory, you need to know the proper amounts of reactants to use and how much of the product will form. Now that we have
SAMPLE PROBLEM 9.2 Calculating Moles of a Reactant

In the chemical reaction of iron and sulfur, how many moles of sulfur are needed to react with 1.42 mol of iron?

$$2\text{Fe}(s) + 3\text{S}(s) \rightarrow \text{Fe}_2\text{S}_3(s)$$

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities (moles).

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.42 mol of Fe</td>
<td>moles of S</td>
<td>mole–mole factor</td>
</tr>
</tbody>
</table>

Equation

$$2\text{Fe}(s) + 3\text{S}(s) \rightarrow \text{Fe}_2\text{S}_3(s)$$

STEP 2 Write a plan to convert the given to the needed quantity (moles).

moles of Fe \[\text{Mole–mole factor}\] moles of S

STEP 3 Use coefficients to write mole–mole factors.

$$\frac{2 \text{ mol of Fe}}{3 \text{ mol S}} = \frac{3 \text{ mol S}}{2 \text{ mol Fe}}$$

STEP 4 Set up the problem to give the needed quantity (moles).

$$1.42 \text{ mol Fe} \times \frac{3 \text{ mol S}}{2 \text{ mol Fe}} = 2.13 \text{ mol of S}$$

STUDY CHECK 9.2

Using the equation in Sample Problem 9.2, calculate the number of moles of iron needed to react with 2.75 mol of sulfur.

ANSWER

1.83 mol of iron

SAMPLE PROBLEM 9.3 Calculating Moles of a Product

Propane gas (C\(_3\)H\(_8\)), a fuel used in camp stoves, soldering torches, and specially equipped automobiles, reacts with oxygen to produce carbon dioxide, water, and energy. How many moles of CO\(_2\) can be produced when 2.25 mol of C\(_3\)H\(_8\) reacts?

$$\text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow \Delta 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g)$$

TRY IT FIRST
**SOLUTION**

**STEP 1** State the given and needed quantities (moles).

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>2.25 mol of C₃H₈</td>
<td>moles of CO₂</td>
<td>mole–mole factor</td>
</tr>
</tbody>
</table>

**Equation**

C₃H₈(g) + 5O₂(g) → △ 3CO₂(g) + 4H₂O(g)

**STEP 2** Write a plan to convert the given to the needed quantity (moles).

moles of C₃H₈ \( \text{Mole–mole factor} \) \( \text{moles of CO}_2 \)

**STEP 3** Use coefficients to write mole–mole factors.

\[ \frac{1 \text{ mol of C}_3\text{H}_8}{3 \text{ mol of CO}_2}, \quad \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \]

**STEP 4** Set up the problem to give the needed quantity (moles).

\[ 2.25 \text{ mol C}_3\text{H}_8 \times \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} = 6.75 \text{ mol of CO}_2 \]

Three SFs Exact Three SFs

**STUDY CHECK 9.3**

Using the equation in Sample Problem 9.3, calculate the moles of oxygen that must react to produce 0.756 mol of water.

**ANSWER**

0.945 mol of O₂

**QUESTIONS AND PROBLEMS**

**9.2 Calculating Moles Using Mole–Mole Factors**

**LEARNING GOAL** Use a mole–mole factor from a balanced chemical equation to calculate the number of moles of another substance in the reaction.

9.3 Write all of the mole–mole factors for each of the following equations:

a. \( \text{2SO}_2(g) + \text{O}_2(g) \rightarrow \text{2SO}_3(g) \)

b. \( \text{4P(s)} + \text{5O}_2(g) \rightarrow \text{2P}_2\text{O}_5(s) \)

9.4 Write all of the mole–mole factors for each of the following equations:

a. \( \text{2Al(s)} + \text{3Cl}_2(g) \rightarrow \text{2AlCl}_3(s) \)

b. \( \text{4HCl(g)} + \text{O}_2(g) \rightarrow \text{2Cl}_2(g) + 2\text{H}_2\text{O(g)} \)

9.5 For the equations in problem 9.3, write the setup with the correct mole–mole factor:

a. moles of SO₃ from the moles of SO₂

b. moles of O₂ needed to react with moles of P

9.6 For the equations in problem 9.4, write the setup with the correct mole–mole factor:

a. moles of AlCl₃ from the moles of Cl₂

b. moles of O₂ needed to react with moles of HCl

9.7 The reaction of hydrogen with oxygen produces water.

\( \text{2H}_2(g) + \text{O}_2(g) \rightarrow \text{2H}_2\text{O(g)} \)

a. How many moles of O₂ are required to react with 2.6 mol of H₂?

b. How many moles of H₂ are required to react with 5.0 mol of O₂?

c. How many moles of H₂O form when 2.5 mol of O₂ reacts?
Mass Calculations for Reactions

**LEARNING GOAL** Given the mass in grams of a substance in a reaction, calculate the mass in grams of another substance in the reaction.

When we have the balanced chemical equation for a reaction, we can use the mass of one of the substances (A) in the reaction to calculate the mass of another substance (B) in the reaction. However, the calculations require us to convert the mass of A to moles of A using the molar mass for A. Then we use the mole–mole factor that links substance A to substance B, which we obtain from the coefficients in the balanced equation. This mole–mole factor (B/A) will convert the moles of A to moles of B. Then the molar mass of B is used to calculate the grams of substance B.

### SAMPLE PROBLEM 9.4 Grams of Product from Grams of Reactant

In the engines of cars and trucks, nitrogen and oxygen from the air react at high temperature to produce nitrogen oxide, a component of smog. Complete the following to help answer the question: How many grams of NO can be produced when 12.5 g of O₂ reacts?

\[ \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \]

a. What molar mass is needed to convert grams of O₂ to moles of O₂?
b. What mole–mole factor is needed to convert moles of O₂ to moles of NO?
c. What molar mass is needed to convert moles of NO to grams of NO?
SOLUTION

a. The molar mass that gives the moles of O\textsubscript{2} is
\[
\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2}
\]

b. The mole–mole factor that gives the moles of NO is
\[
\frac{2 \text{ mol NO}}{1 \text{ mol O}_2}
\]

c. The molar mass that gives the grams of NO is
\[
\frac{30.01 \text{ g NO}}{1 \text{ mol NO}}
\]

STUDY CHECK 9.4
Using the equation in Sample Problem 9.4, what molar mass is needed to convert grams of N\textsubscript{2} to moles of N\textsubscript{2}?

ANSWER
\[
\begin{align*}
1 \text{ mol N}_2 &= 28.02 \text{ g N}_2 \\
\end{align*}
\]

SAMPLE PROBLEM 9.5 Grams of Product

When acetylene, C\textsubscript{2}H\textsubscript{2}, burns in oxygen, high temperatures are produced that are used for welding metals.

\[
2\text{C}_2\text{H}_2(g) + 5\text{O}_2(g) \xrightarrow{\Delta} 4\text{CO}_2(g) + 2\text{H}_2\text{O}(g)
\]

How many grams of CO\textsubscript{2} are produced when 54.6 g of C\textsubscript{2}H\textsubscript{2} is burned?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities (grams).

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>54.6 g of C\textsubscript{2}H\textsubscript{2}</td>
<td>grams of CO\textsubscript{2}</td>
<td>molar masses, mole–mole factor</td>
</tr>
</tbody>
</table>

Equation

\[
2\text{C}_2\text{H}_2(g) + 5\text{O}_2(g) \xrightarrow{\Delta} 4\text{CO}_2(g) + 2\text{H}_2\text{O}(g)
\]

STEP 2 Write a plan to convert the given to the needed quantity (grams).

<table>
<thead>
<tr>
<th>grams of C\textsubscript{2}H\textsubscript{2}</th>
<th>Molar mass</th>
<th>moles of C\textsubscript{2}H\textsubscript{2}</th>
<th>Mole–mole factor</th>
<th>moles of CO\textsubscript{2}</th>
<th>Molar mass</th>
<th>grams of CO\textsubscript{2}</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mol of C\textsubscript{2}H\textsubscript{2} = 26.04 g of C\textsubscript{2}H\textsubscript{2}</td>
<td>26.04 g C\textsubscript{2}H\textsubscript{2}</td>
<td>1 mol C\textsubscript{2}H\textsubscript{2}</td>
<td>26.04 g C\textsubscript{2}H\textsubscript{2}</td>
<td>2 mol of C\textsubscript{2}H\textsubscript{2} = 4 mol of CO\textsubscript{2}</td>
<td>44.01 g CO\textsubscript{2}</td>
<td>1 mol of CO\textsubscript{2} = 44.01 g of CO\textsubscript{2}</td>
</tr>
</tbody>
</table>

STEP 3 Use coefficients to write mole–mole factors; write molar masses.

A mixture of acetylene and oxygen undergoes combustion during the welding of metals.
Mass Calculations for Reactions

STEP 4 Set up the problem to give the needed quantity (grams).

\[
54.6 \text{ g } C_2H_2 \times \frac{1 \text{ mol } C_2H_2}{26.04 \text{ g } C_2H_2} \times \frac{4 \text{ mol } CO_2}{2 \text{ mol } C_2H_2} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 185 \text{ g of } CO_2
\]

STUDY CHECK 9.5

Using the equation in Sample Problem 9.5, calculate the grams of CO₂ that can be produced when 25.0 g of O₂ reacts.

ANSWER

27.5 g of CO₂

SAMPLE PROBLEM 9.6 Grams of Reactant

The fuel heptane (C₇H₁₆) is designated as the zero point in the octane rating of gasoline. Heptane is an undesirable compound in gasoline because it burns rapidly and causes engine knocking. How many grams of O₂ are required to react with 22.5 g of C₇H₁₆?

\[
C_7H_{16}(l) + 11O_2(g) \rightarrow 7CO_2(g) + 8H₂O(g)
\]

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities (grams).

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>Equation</td>
<td>22.5 g of C₇H₁₆</td>
<td>grams of O₂</td>
<td>molar masses, mole–mole factor</td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to convert the given to the needed quantity (grams).

<table>
<thead>
<tr>
<th>Given</th>
<th>Plan</th>
<th>Need</th>
</tr>
</thead>
<tbody>
<tr>
<td>grams of C₇H₁₆</td>
<td>Molar mass</td>
<td>moles of C₇H₁₆</td>
</tr>
<tr>
<td>Mole–mole factor</td>
<td></td>
<td>moles of O₂</td>
</tr>
<tr>
<td>Molar mass</td>
<td></td>
<td>grams of O₂</td>
</tr>
</tbody>
</table>

STEP 3 Use coefficients to write mole–mole factors; write molar masses.

\[
1 \text{ mol of C}_7\text{H}_{16} = \frac{100.2 \text{ g of C}_7\text{H}_{16}}{1 \text{ mol C}_7\text{H}_{16}}
\]

\[
1 \text{ mol of C}_7\text{H}_{16} = \frac{11 \text{ mol O}_2}{1 \text{ mol C}_7\text{H}_{16}}
\]

\[
1 \text{ mol of } O_2 = \frac{32.00 \text{ g of O}_2}{1 \text{ mol O}_2}
\]

STEP 4 Set up the problem to give the needed quantity (grams).

\[
22.5 \text{ g } C_7H_{16} \times \frac{1 \text{ mol } C_7H_{16}}{100.2 \text{ g } C_7H_{16}} \times \frac{11 \text{ mol } O_2}{1 \text{ mol } C_7H_{16}} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 79.1 \text{ g of } O_2
\]
CHAPTER 9 Chemical Quantities in Reactions

STUDY CHECK 9.6
Using the equation in Sample Problem 9.6, calculate the grams of C\(_7\)H\(_{16}\) that are needed to produce 15.0 g of H\(_2\)O.

ANSWER
10.4 g of C\(_7\)H\(_{16}\)

QUESTIONS AND PROBLEMS

9.3 Mass Calculations for Reactions

LEARNING GOAL Given the mass in grams of a substance in a reaction, calculate the mass in grams of another substance in the reaction.

9.11 Sodium reacts with oxygen to produce sodium oxide.

\[ 4\text{Na(s)} + \text{O}_2(g) \rightarrow 2\text{Na}_2\text{O(s)} \]

a. How many grams of Na\(_2\)O are produced when 57.5 g of Na reacts?
b. If you have 18.0 g of Na, how many grams of O\(_2\) are required for reaction?
c. How many grams of O\(_2\) are needed in a reaction that produces 75.0 g of Na\(_2\)O?

9.12 Nitrogen gas reacts with hydrogen gas to produce ammonia.

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \]

a. If you have 3.64 g of H\(_2\), how many grams of NH\(_3\) can be produced?
b. How many grams of H\(_2\) are needed to react with 2.80 g of N\(_2\)?
c. How many grams of NH\(_3\) can be produced from 12.0 g of H\(_2\)?

9.13 Ammonia and oxygen react to form nitrogen and water.

\[ 4\text{NH}_3(g) + 3\text{O}_2(g) \rightarrow 2\text{N}_2(g) + 6\text{H}_2\text{O}(g) \]

a. How many grams of O\(_2\) are needed to react with 13.6 g of NH\(_3\)?
b. How many grams of N\(_2\) can be produced when 6.50 g of O\(_2\) reacts?
c. How many grams of H\(_2\)O are formed from the reaction of 34.0 g of NH\(_3\)?

9.14 Iron(III) oxide reacts with carbon to give iron and carbon monoxide.

\[ \text{Fe}_2\text{O}_3(s) + 3\text{C(s)} \rightarrow 2\text{Fe(s)} + 3\text{CO(g)} \]

a. How many grams of C are required to react with 16.5 g of Fe\(_2\)O\(_3\)?
b. How many grams of CO are produced when 36.0 g of C reacts?
c. How many grams of Fe can be produced when 6.00 g of Fe\(_2\)O\(_3\) reacts?

9.15 Nitrogen dioxide and water react to produce nitric acid, HNO\(_3\), and nitrogen oxide.

\[ 3\text{NO}_2(g) + \text{H}_2\text{O(l)} \rightarrow 2\text{HNO}_3(aq) + \text{NO(g)} \]

a. How many grams of H\(_2\)O are required to react with 28.0 g of NO\(_2\)?
b. How many grams of NO are produced from 15.8 g of H\(_2\)O?
c. How many grams of HNO\(_3\) are produced from 8.25 g of NO\(_2\)?

9.16 Calcium cyanamide, CaCN\(_2\), reacts with water to form calcium carbonate and ammonia.

\[ \text{CaCN}_2(s) + 3\text{H}_2\text{O(l)} \rightarrow \text{CaCO}_3(s) + 2\text{NH}_3(g) \]

a. How many grams of H\(_2\)O are needed to react with 75.0 g of CaCN\(_2\)?
b. How many grams of NH\(_3\) are produced from 5.24 g of CaCN\(_2\)?
c. How many grams of CaCO\(_3\) form if 155 g of H\(_2\)O reacts?

9.17 When solid lead(II) sulfide reacts with oxygen gas, the products are solid lead(II) oxide and sulfur dioxide gas.

a. Write the balanced equation for the reaction.
b. How many grams of oxygen are required to react with 29.9 g of lead(II) sulfide?
c. How many grams of sulfur dioxide can be produced when 65.0 g of lead(II) sulfide reacts?
d. How many grams of lead(II) sulfide are used to produce 128 g of lead(II) oxide?

9.18 When the gases dihydrogen sulfide and oxygen react, they form the gases sulfur dioxide and water vapor.

a. Write the balanced equation for the reaction.
b. How many grams of oxygen are required to react with 2.50 g of dihydrogen sulfide?
c. How many grams of sulfur dioxide can be produced when 38.5 g of oxygen reacts?
d. How many grams of oxygen are required to produce 55.8 g of water vapor?

9.4 Limiting Reactants

LEARNING GOAL Identify a limiting reactant when given the quantities of two reactants; calculate the amount of product formed from the limiting reactant.

When we make peanut butter sandwiches for lunch, we need 2 slices of bread and 1 tablespoon of peanut butter for each sandwich. As an equation, we could write:

\[ 2 \text{ slices of bread} + 1 \text{ tablespoon of peanut butter} \rightarrow 1 \text{ peanut butter sandwich} \]
If we have 8 slices of bread and a full jar of peanut butter, we will run out of bread after we make 4 peanut butter sandwiches. We cannot make any more sandwiches once the bread is used up, even though there is a lot of peanut butter left in the jar. The number of slices of bread has limited the number of sandwiches we can make.

On a different day, we might have 8 slices of bread but only a tablespoon of peanut butter left in the peanut butter jar. We will run out of peanut butter after we make just 1 peanut butter sandwich and have 6 slices of bread left over. The small amount of peanut butter available has limited the number of sandwiches we can make.

The reactant that is completely used up is the limiting reactant. The reactant that does not completely react and is left over is called the excess reactant.

<table>
<thead>
<tr>
<th>Bread</th>
<th>Peanut</th>
<th>Sandwiches</th>
<th>Limiting Reactant</th>
<th>Excess Reactant</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 loaf (20 slices)</td>
<td>1 tablespoon</td>
<td>1</td>
<td>peanut butter</td>
<td>bread</td>
</tr>
<tr>
<td>4 slices</td>
<td>1 full jar</td>
<td>2</td>
<td>bread</td>
<td>peanut butter</td>
</tr>
<tr>
<td>8 slices</td>
<td>1 full jar</td>
<td>4</td>
<td>bread</td>
<td>peanut butter</td>
</tr>
</tbody>
</table>

Calculating Moles of Product from a Limiting Reactant

In a similar way, the reactants in a chemical reaction do not always combine in quantities that allow each to be used up at exactly the same time. In many reactions, there is a limiting reactant that determines the amount of product that can be formed. When we know the quantities of the reactants of a chemical reaction, we calculate the amount of product that is possible from each reactant if it were completely consumed. We are looking for the limiting reactant, which is the one that runs out first, producing the smaller amount of product.

**SAMPLE PROBLEM 9.7 Moles of Product from Limiting Reactant**

Carbon monoxide and hydrogen are used to produce methanol (CH₄O). The balanced chemical reaction is

\[
\text{CO}(g) + 2\text{H}_2(g) \rightarrow \text{CH}_4\text{O}(g)
\]

If 3.00 mol of CO and 5.00 mol of H₂ are the initial reactants, what is the limiting reactant and how many moles of methanol can be produced?
SOLUTION

**STEP 1** State the given and needed quantity (moles).

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>3.00 mol of CO</td>
<td>moles of CH₄O produced, moles of CH₄O produced,</td>
<td>mole–mole factors</td>
</tr>
<tr>
<td>5.00 mol of H₂</td>
<td>limiting reactant</td>
<td></td>
</tr>
</tbody>
</table>

**Equation**

\[ \text{CO(g) + 2H}_2\text{(g) } \rightarrow \text{CH}_4\text{O(g)} \]

**STEP 2** Write a plan to convert the quantity (moles) of each reactant to quantity (moles) of product.

- moles of CO: Mole–mole factor
- moles of CH₄O

- moles of H₂: Mole–mole factor
- moles of CH₄O

**STEP 3** Write the mole–mole factors.

\[
\begin{align*}
1 \text{ mol CO} & = 1 \text{ mol CH}_4\text{O} \\
1 \text{ mol CO} & = 1 \text{ mol CH}_4\text{O} \\
2 \text{ mol H}_2 & = 1 \text{ mol CH}_4\text{O} \\
1 \text{ mol CH}_4\text{O} & = 2 \text{ mol H}_2
\end{align*}
\]

**STEP 4** Calculate the quantity (moles) of product from each reactant and select the smaller quantity (moles) as the limiting reactant.

**Moles of CH₄O (product) from CO:**

- Exact
- Three SFs

\[
3.00 \text{ mol CO} \times \frac{1 \text{ mol CH}_4\text{O}}{1 \text{ mol CO}} = 3.00 \text{ mol of CH}_4\text{O}
\]

**Moles of CH₄O (product) from H₂:**

- Exact
- Three SFs
- Smaller amount of product

\[
5.00 \text{ mol H}_2 \times \frac{1 \text{ mol CH}_4\text{O}}{2 \text{ mol H}_2} = 2.50 \text{ mol of CH}_4\text{O}
\]

The smaller amount, 2.50 mol of CH₄O, is the maximum amount of methanol that can be produced from the limiting reactant, H₂, because it is completely consumed.

**STUDY CHECK 9.7**

If the initial mixture of reactants for Sample Problem 9.7 contains 4.00 mol of CO and 4.00 mol of H₂, what is the limiting reactant and how many moles of methanol can be produced?

**ANSWER**

H₂ is the limiting reactant; 2.00 mol of methanol can be produced.
Calculating Mass of Product from a Limiting Reactant

The quantities of the reactants can also be given in grams. The calculations to identify the limiting reactant are the same as before, but the grams of each reactant must first be converted to moles, then to moles of product, and finally to grams of product. Then select the smaller mass of product, which is from complete use of the limiting reactant. This calculation is shown in Sample Problem 9.8.

**SAMPLE PROBLEM 9.8 Grams of Product from a Limiting Reactant**

When silicon dioxide (sand) and carbon are heated, the products are silicon carbide, SiC, and carbon monoxide. Silicon carbide is a ceramic material that tolerates extreme temperatures and is used as an abrasive and in the brake discs of sports cars. How many grams of CO are formed from 70.0 g of SiO₂ and 50.0 g of C?

\[ \text{SiO}_2(s) + 3\text{C}(s) \xrightarrow{\Delta} \text{SiC}(s) + 2\text{CO}(g) \]

**SOLUTION**

**STEP 1** State the given and needed quantity (grams).

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>70.0 g of SiO₂, 50.0 g of C</td>
<td>grams of CO</td>
<td>molar masses, limiting reactant, mole–mole factor</td>
</tr>
<tr>
<td>Equation</td>
<td>( \text{SiO}_2(s) + 3\text{C}(s) \xrightarrow{\Delta} \text{SiC}(s) + 2\text{CO}(g) )</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to convert the quantity (grams) of each reactant to quantity (grams) of product.

- grams of SiO₂ \( \rightarrow \) moles of SiO₂ \( \rightarrow \) moles of CO \( \rightarrow \) grams of CO
- grams of C \( \rightarrow \) moles of C \( \rightarrow \) moles of CO \( \rightarrow \) grams of CO

**STEP 3** Write the mole–mole factors and molar masses.

**Molar masses:**

\[
\frac{1 \text{ mol of \text{SiO}_2}}{60.09 \text{ g \text{SiO}_2}} \quad \frac{1 \text{ mol \text{CO}}}{12.01 \text{ g \text{CO}}} \quad \frac{1 \text{ mol \text{CO}}}{28.01 \text{ g \text{CO}}}
\]

**Mole–mole factors:**

\[
\frac{2 \text{ mol of \text{CO}}}{1 \text{ mol \text{SiO}_2}} \quad \frac{3 \text{ mol of \text{C}}}{2 \text{ mol \text{CO}}}
\]
Calculate the quantity (grams) of product from each reactant and select the smaller quantity (grams) as the limiting reactant.

**Grams of CO (product) from SiO₂:**

<table>
<thead>
<tr>
<th>Three SFs</th>
<th>Exact</th>
<th>Four SFs</th>
<th>Three SFs</th>
</tr>
</thead>
<tbody>
<tr>
<td>70.0 g SiO₂</td>
<td>$\frac{1 \text{ mol SiO₂}}{60.09 \text{ g SiO₂}}$</td>
<td>$\frac{2 \text{ mol CO}}{1 \text{ mol SiO₂}}$</td>
<td>$\frac{28.01 \text{ g CO}}{1 \text{ mol CO}}$</td>
</tr>
<tr>
<td>Limiting reactant</td>
<td>Four SFs</td>
<td>Exact</td>
<td>Exact</td>
</tr>
<tr>
<td>Smaller amount of product</td>
<td>$= 65.3 \text{ g of CO}$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Grams of CO (product) from C:**

<table>
<thead>
<tr>
<th>Three SFs</th>
<th>Exact</th>
<th>Four SFs</th>
<th>Three SFs</th>
</tr>
</thead>
<tbody>
<tr>
<td>50.0 g C</td>
<td>$\frac{1 \text{ mol C}}{12.01 \text{ g C}}$</td>
<td>$\frac{2 \text{ mol CO}}{3 \text{ mol C}}$</td>
<td>$\frac{28.01 \text{ g CO}}{1 \text{ mol CO}}$</td>
</tr>
<tr>
<td>Four SFs</td>
<td>Exact</td>
<td>Exact</td>
<td></td>
</tr>
<tr>
<td>$= 77.7 \text{ g of CO}$</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

The smaller amount, 65.3 g of CO, is the most CO that can be produced. This also means that the SiO₂ is the limiting reactant.

**STUDY CHECK 9.8**

Hydrogen sulfide burns with oxygen to give sulfur dioxide and water. How many grams of sulfur dioxide are formed from the reaction of 8.52 g of H₂S and 9.60 g of O₂?

\[2 \text{H₂S(g)} + 3 \text{O₂(g)} \xrightarrow{\Delta} 2 \text{SO₂(g)} + 2 \text{H₂O(g)}\]

**ANSWER**

12.8 g of SO₂
For each of the following reactions, calculate the grams of indicated product when 25.0 g of the first reactant and 40.0 g of the second reactant are used:

- **a.** \(2\text{SO}_3(g) + \text{O}_2(g) \rightarrow 2\text{SO}_2(g)\) (SO

- **b.** \(3\text{Fe}(s) + 4\text{H}_2\text{O}(l) \rightarrow \text{Fe}_2\text{O}_3(s) + 4\text{H}_2(g)\) (Fe)

- **c.** \(\text{C}_2\text{H}_6(l) + 11\text{O}_2(g) \rightarrow 7\text{CO}_2(g) + 8\text{H}_2\text{O}(g)\) (CO

For each of the following reactions, calculate the grams of indicated product when 15.0 g of the first reactant and 10.0 g of the second reactant are used:

- **a.** \(4\text{Li}(s) + \text{O}_2(g) \rightarrow 2\text{Li}_2\text{O}(s)\) (Li

- **b.** \(\text{Fe}_2\text{O}_3(s) + 3\text{H}_2(g) \rightarrow 2\text{Fe}(s) + 3\text{H}_2\text{O}(l)\) (Fe)

- **c.** \(\text{Al}_2\text{S}_3(s) + 6\text{H}_2\text{O}(l) \rightarrow 2\text{Al(OH)}_3(aq) + 3\text{H}_2\text{S}(g)\) (H

### 9.5 Percent Yield

**LEARNING GOAL** Given the actual quantity of product, determine the percent yield for a reaction.

In our problems up to now, we assumed that all of the reactants changed completely to product. Thus, we have calculated the amount of product as the maximum quantity possible, or 100%. While this would be an ideal situation, it does not usually happen. As we carry out a reaction and transfer products from one container to another, some product is usually lost. In the lab as well as commercially, the starting materials may not be completely pure, and side reactions may use some of the reactants to give unwanted products. Thus, 100% of the desired product is not actually obtained.

When we do a chemical reaction in the laboratory, we measure our specific quantities of the reactants. We calculate the **theoretical yield** for the reaction, which is the amount of product (100%) we would expect if all the reactants were converted to the desired product. When the reaction ends, we collect and measure the mass of the product, which is the **actual yield** for the product. Because some product is usually lost, the actual yield is less than the theoretical yield. Using the actual yield and the theoretical yield for a product, we can calculate the **percent yield**.

\[
\text{Percent yield (\%) = } \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%
\]

### SAMPLE PROBLEM 9.9 Calculating Percent Yield

On a space shuttle, LiOH is used to absorb exhaled CO\(_2\) from breathing air to form LiHCO\(_3\).

\[
\text{LiOH}(s) + \text{CO}_2(g) \rightarrow \text{LiHCO}_3(s)
\]

What is the percent yield of LiHCO\(_3\) for the reaction if 50.0 g of LiOH gives 72.8 g of LiHCO\(_3\)?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>50.0 g of LiOH,</td>
<td>percent yield of</td>
<td>molar masses, mole–</td>
</tr>
<tr>
<td>72.8 g of LiHCO(_3)</td>
<td>LiHCO(_3)</td>
<td>mole factor, percent</td>
</tr>
<tr>
<td></td>
<td></td>
<td>yield expression</td>
</tr>
</tbody>
</table>

**Equation**

\[
\text{LiOH}(s) + \text{CO}_2(g) \rightarrow \text{LiHCO}_3(s)
\]
**Guide to Calculating Percent Yield**

**STEP 1**
State the given and needed quantities.

**STEP 2**
Write a plan to calculate the theoretical yield and the percent yield.

**STEP 3**
Write the molar masses and the mole–mole factor from the balanced equation.

**STEP 4**
Calculate the percent yield by dividing the actual yield (given) by the theoretical yield and multiplying the result by 100%.

---

### Calculation of theoretical yield:

**Grams of LiOH**

<table>
<thead>
<tr>
<th>Molar mass</th>
<th>moles of LiOH</th>
<th>Mole–mole factor</th>
<th>moles of LiHCO₃</th>
<th>Molar mass</th>
<th>grams of LiHCO₃</th>
</tr>
</thead>
</table>

#### Calculation of percent yield:

Percent yield (%) = \( \frac{\text{actual yield of LiHCO}_3}{\text{theoretical yield of LiHCO}_3} \times 100\% \)

**STEP 3**
Write the molar masses and the mole–mole factor from the balanced equation.

**Molar masses:**

- 1 mol of LiOH = 23.95 g of LiOH
- 1 mol of LiHCO₃ = 67.96 g of LiHCO₃

**Mole–mole factor:**

- 1 mol of LiHCO₃ = 1 mol of LiOH
- 1 mol LiHCO₃ and 1 mol LiOH

**STEP 4**
Calculate the percent yield by dividing the actual yield (given) by the theoretical yield and multiplying the result by 100%.

**Calculation of theoretical yield:**

\[
50.0 \text{ g LiOH} \times \frac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}} \times \frac{1 \text{ mol LiHCO}_3}{1 \text{ mol LiOH}} \times \frac{67.96 \text{ g LiHCO}_3}{1 \text{ mol LiHCO}_3} = 142 \text{ g of LiHCO}_3
\]

**Calculation of percent yield:**

\[
\frac{72.8 \text{ g LiHCO}_3}{142 \text{ g LiHCO}_3} \times 100\% = 51.3\%
\]

A percent yield of 51.3% means that 72.8 g of the theoretical amount of 142 g of LiHCO₃ was actually produced by the reaction.

**STUDY CHECK 9.9**

For the reaction in Sample Problem 9.9, what is the percent yield of LiHCO₃ if 8.00 g of CO₂ produces 10.5 g of LiHCO₃?

**ANSWER**

84.7%
Questions and Problems

9.5 Percent Yield

Learning Goal: Given the actual quantity of product, determine the percent yield for a reaction.

9.27 Carbon disulfide is produced by the reaction of carbon and sulfur dioxide.

\[ 5C(s) + 2SO_2(g) \rightarrow CS_2(g) + 4CO(g) \]

a. What is the percent yield of carbon disulfide if the reaction of 40.0 g of carbon produces 36.0 g of carbon disulfide?
b. What is the percent yield of carbon disulfide if the reaction of 32.0 g of sulfur dioxide produces 12.0 g of carbon disulfide?

9.28 Iron(III) oxide reacts with carbon monoxide to produce iron and carbon dioxide.

\[ Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g) \]

a. What is the percent yield of iron if the reaction of 65.0 g of iron(III) oxide produces 15.0 g of iron?
b. What is the percent yield of carbon dioxide if the reaction of 75.0 g of carbon monoxide produces 85.0 g of carbon dioxide?

9.29 Aluminum reacts with oxygen to produce aluminum oxide.

\[ 4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s) \]

Calculate the mass of Al_2O_3 that can be produced if the reaction of 50.0 g of aluminum and sufficient oxygen has a 75.0% yield.

9.30 Propane (C_3H_8) burns in oxygen to produce carbon dioxide and water.

\[ C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) \]

Calculate the mass of CO_2 that can be produced if the reaction of 45.0 g of propane and sufficient oxygen has a 60.0% yield.

9.31 When 30.0 g of carbon is heated with silicon dioxide, 28.2 g of carbon monoxide is produced. What is the percent yield of carbon monoxide for this reaction?

\[ SiO_2(s) + 3C(s) \rightarrow SiC(s) + 2CO(g) \]

9.32 When 56.6 g of calcium is reacted with nitrogen gas, 32.4 g of calcium nitride is produced. What is the percent yield of calcium nitride for this reaction?

\[ 3Ca(s) + N_2(g) \rightarrow Ca_3N_2(s) \]

9.6 Energy in Chemical Reactions

Learning Goal: Given the heat of reaction (enthalpy change), calculate the loss or gain of heat for an exothermic or endothermic reaction.

Almost every chemical reaction involves a loss or gain of energy. To discuss energy change (enthalpy change) for a reaction, we look at the energy that is absorbed or lost during a chemical reaction.

Energy Units for Chemical Reactions

The SI unit for energy is the joule (J). Often, the unit of kilojoules (kJ) is used to show the energy change in a reaction.

1 kilojoule (kJ) = 1000 joules (J)

Heat of Reaction (Enthalpy Change)

The heat of reaction is the amount of heat absorbed or released during a reaction that takes place at constant pressure. A change of energy occurs as reactants interact, bonds break apart, and products form. We determine a heat of reaction, symbol ΔH, as the difference in the energy of the products and the reactants.

\[ \Delta H = H_{\text{products}} - H_{\text{reactants}} \]

Exothermic Reactions

In an exothermic reaction (exo means “out”), the energy of the products is lower than that of the reactants. This means that heat is released along with the products that form. Let us look at the equation for the exothermic reaction in which 185 kJ of heat is released when 1 mol of hydrogen and 1 mol of chlorine react to form 2 mol of hydrogen chloride. For an
Exothermic reaction, the heat of reaction can be written as one of the products. It can also be written as a $\Delta H$ value with a negative sign ($-$).

\[
\text{H}_2 (g) + \text{Cl}_2 (g) \rightarrow 2\text{HCl}(g) + 185\text{ kJ} \quad \Delta H = -185\text{ kJ}
\]

**Endothermic Reactions**

In an endothermic reaction (*endo* means “within”), the energy of the products is higher than that of the reactants. Heat is required to convert the reactants to products. Let us look at the equation for the endothermic reaction in which 180 kJ of heat is needed to convert 1 mol of nitrogen and 1 mol of oxygen to 2 mol of nitrogen oxide. For an endothermic reaction, the heat of reaction can be written as one of the reactants. It can also be written as a $\Delta H$ value with a positive sign ($+$).

\[
\text{N}_2(g) + \text{O}_2(g) + 180\text{ kJ} \rightarrow 2\text{NO}(g) \quad \Delta H = +180\text{ kJ}
\]

**SAMPLE PROBLEM 9.10 Exothermic and Endothermic Reactions**

In the reaction of 1 mol of solid carbon with oxygen gas, the energy of the carbon dioxide gas produced is 393 kJ less than that of the reactants.

a. Is the reaction exothermic or endothermic?
b. Write the balanced chemical equation including the heat of the reaction.
c. What is the value, in kilojoules, for the heat of reaction?
Calculations of Heat in Reactions

The value of $\Delta H$ refers to the heat change in kilojoules for each substance in the balanced equation for the reaction. Consider the following decomposition reaction:

$$2\text{H}_2\text{O}(l) \longrightarrow 2\text{H}_2(g) + \text{O}_2(g) \quad \Delta H = +572 \text{ kJ}$$

For this reaction, 572 kJ are absorbed by 2 mol of $\text{H}_2\text{O}$ to produce 2 mol of $\text{H}_2$ and 1 mol of $\text{O}_2$. We can write heat conversion factors for each substance in this reaction as follows:

$$\frac{+572 \text{ kJ}}{2 \text{ mol } \text{H}_2\text{O}} \quad \frac{+572 \text{ kJ}}{2 \text{ mol } \text{H}_2} \quad \frac{+572 \text{ kJ}}{1 \text{ mol } \text{O}_2}$$

Suppose in this reaction that 9.00 g of $\text{H}_2\text{O}$ undergoes reaction. We can calculate the quantity of heat absorbed as

$$9.00 \text{ g } \text{H}_2\text{O} \times \frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g } \text{H}_2\text{O}} \times \frac{+572 \text{ kJ}}{2 \text{ mol } \text{H}_2\text{O}} = +143 \text{ kJ}$$

**SAMPLE PROBLEM 9.11 Calculating Heat in a Reaction**

How much heat, in kilojoules, is released when nitrogen and hydrogen react to form 50.0 g of ammonia?

$$\text{N}_2(g) + 3\text{H}_2(g) \longrightarrow 2\text{NH}_3(g) \quad \Delta H = -92.2 \text{ kJ}$$

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>50.0 g of $\text{NH}_3$, $\Delta H = -92.2$ kJ</td>
<td>heat released, in kilojoules</td>
<td>molar mass, heat conversion factor</td>
</tr>
</tbody>
</table>

**Equation**

$$\text{N}_2(g) + 3\text{H}_2(g) \longrightarrow 2\text{NH}_3(g)$$
**Guide to Calculations Using the Heat of Reaction (ΔH)**

**STEP 1**
State the given and needed quantities.

**STEP 2**
Write a plan using the heat of reaction and any molar mass needed.

**STEP 3**
Write the conversion factors including heat of reaction.

**STEP 4**
Set up the problem to calculate the heat.

---

### CHEMISTRY LINK TO HEALTH

**Cold Packs and Hot Packs**

In a hospital, at a first-aid station, or at an athletic event, an instant cold pack may be used to reduce swelling from an injury, remove heat from inflammation, or decrease capillary size to lessen the effect of hemorrhaging. Inside the plastic container of a cold pack, there is a compartment containing solid ammonium nitrate (NH₄NO₃) that is separated from a compartment containing water. The pack is activated when it is hit or squeezed hard enough to break the walls between the compartments and cause the ammonium nitrate to mix with the water (shown as H₂O over the reaction arrow). In an endothermic process, 1 mol of NH₄NO₃ that dissolves absorbs 26 kJ of heat. The temperature drops to about 4 to 5 °C to give a cold pack that is ready to use.

**Endothermic Reaction in a Cold Pack**

\[
\text{NH}_4\text{NO}_3(s) + 26 \text{kJ} \xrightarrow{\text{H}_2\text{O}} \text{NH}_4\text{NO}_3(aq)
\]

Hot packs are used to relax muscles, lessen aches and cramps, and increase circulation by expanding capillary size. Constructed in the same way as cold packs, a hot pack contains a salt such as CaCl₂. When 1 mol of CaCl₂ dissolves in water, 82 kJ are released as heat. The temperature increases as much as 66 °C to give a hot pack that is ready to use.

**Exothermic Reaction in a Hot Pack**

\[
\text{CaCl}_2(s) \xrightarrow{\text{H}_2\text{O}} \text{CaCl}_2(aq) + 82 \text{kJ}
\]

---

### Example Problem

To determine the heat of reaction for the decomposition of mercury(II) oxide, we can follow these steps:

**STEP 2**
Write a plan using the heat of reaction and any molar mass needed.

**STEP 3**
Write the conversion factors including heat of reaction.

<table>
<thead>
<tr>
<th>Moles of NH₃</th>
<th>Heat of reaction</th>
<th>Kilojoules</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mol NH₃</td>
<td>17.03 g NH₃</td>
<td>92.2 kJ</td>
</tr>
</tbody>
</table>

**STEP 4**
Set up the problem to calculate the heat.

\[
50.0 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \times \frac{-92.2 \text{ kJ}}{2 \text{ mol NH}_3} = -135 \text{ kJ}
\]
Hess’s Law

According to Hess’s law, heat can be absorbed or released in a single chemical reaction or in several steps. When there are two or more steps in the reaction, the overall enthalpy change is the sum of the enthalpy changes of those steps, provided they all occur at the same temperature.

Steps in solving problems involving Hess’s law:

1. If you reverse a chemical equation, you must also reverse the sign of $\Delta H$.
2. If a chemical equation is multiplied by some factor, then $\Delta H$ must be multiplied by the same factor.

We can see how the enthalpy change for a specific reaction is the sum of two or more reactions in Sample Problem 9.12.

**Sample Problem 9.12 Hess’s Law and Calculating Heat of Reaction**

Calculate the $\Delta H$ value for the following reaction using equations 1 and 2:

$$
\text{C(s) + 2H}_2\text{O(g) \rightarrow CO}_2\text{(g) + 2H}_2\text{(g)}
$$

Equation 1  \[ \text{C(s) + O}_2\text{(g) \rightarrow CO}_2\text{(g) \quad \Delta H = -304 \text{ kJ}} \]

Equation 2  \[ \text{H}_2\text{(g) + 1/2 O}_2\text{(g) \rightarrow H}_2\text{O(g) \quad \Delta H = -242 \text{ kJ}} \]

**Try it First**

**Solution**

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>overall reaction, equations 1, 2</td>
<td>heat of reaction</td>
<td>combine equations, total heats of reactions</td>
<td></td>
</tr>
</tbody>
</table>

By combining, rearranging, and multiplying the substances including the $\Delta H$ in equations 1 and 2, we can obtain the overall equation for the reaction. Then the sum of the $\Delta H$ values involved in each will give the $\Delta H$ value for the reaction.

**Step 1** Arrange the given equations to place reactants on the left and products on the right.

$$
\text{C(s) + O}_2\text{(g) \rightarrow CO}_2\text{(g) \quad \Delta H = -304 \text{ kJ}}
$$

$$
\text{H}_2\text{O(g) \rightarrow H}_2\text{(g) + 1/2 O}_2\text{(g) \quad \Delta H = +242 \text{ kJ}}
$$

**Step 2** If an equation is multiplied to balance coefficients, multiply the $\Delta H$ by the same number. The second equation must be multiplied by 2 to give 2H$_2$O(g) on the reactant side.

$$
2 \times \text{[H}_2\text{O(g) \rightarrow H}_2\text{(g) + 1/2 O}_2\text{(g) \quad \Delta H = +242 \text{ kJ}}]
$$

$$
2\text{H}_2\text{O(g) \rightarrow 2H}_2\text{(g) + O}_2\text{(g) \quad \Delta H = +484 \text{ kJ}}
$$

**Step 3** Combine the equations and cancel any substances that are common to both sides. Add the $\Delta H$s.

$$
\text{C(s) + O}_2\text{(g) \rightarrow CO}_2\text{(g) \quad \Delta H = -304 \text{ kJ}}
$$

$$
2\text{H}_2\text{O(g) \rightarrow 2H}_2\text{(g) + O}_2\text{(g) \quad \Delta H = +484 \text{ kJ}}
$$

$$
\text{C(s) + 2H}_2\text{O(g) \rightarrow CO}_2\text{(g) + 2H}_2\text{(g) \quad \Delta H = +180 \text{ kJ}}
$$

**Study Check 9.12**

Calculate the $\Delta H$ value for the following reaction using equations 1 and 2:

$$
\text{2NO(g) + O}_2\text{(g) \rightarrow N}_2\text{O}_4\text{(g)}
$$

Equation 1  \[ \text{N}_2\text{O}_4\text{(g) \rightarrow 2NO}_2\text{(g) \quad \Delta H = +57.2 \text{ kJ}} \]

Equation 2  \[ \text{NO(g) + 1/2 O}_2\text{(g) \rightarrow NO}_2\text{(g) \quad \Delta H = -57.0 \text{ kJ}} \]

**Answer**

$\Delta H = -171.2 \text{ kJ}$
9.33 In an exothermic reaction, is the energy of the products higher or lower than that of the reactants?

9.34 In an endothermic reaction, is the energy of the products higher or lower than that of the reactants?

9.35 Classify each of the following as exothermic or endothermic:
   a. A reaction releases 550 kJ.
   b. The energy level of the products is higher than that of the reactants.
   c. The metabolism of glucose in the body provides energy.

9.36 Classify each of the following as exothermic or endothermic:
   a. The energy level of the products is lower than that of the reactants.
   b. In the body, the synthesis of proteins requires energy.
   c. A reaction absorbs 125 kJ.

9.37 Classify each of the following as exothermic or endothermic and give the \( \Delta H \) for each:
   a. \( \text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g) \) \( \Delta H = +802 \text{ kJ} \)
   b. \( \text{Ca(OH)}_2(s) \rightarrow \text{CaO}(s) + \text{H}_2\text{O}(l) \) \( \Delta H = +65.3 \text{ kJ} \)
   c. \( \text{2Al}(s) + \text{Fe}_2\text{O}_3(s) \rightarrow \text{Al}_2\text{O}_3(s) + 2\text{Fe}(l) \) \( \Delta H = +850 \text{ kJ} \)

9.38 Classify each of the following as exothermic or endothermic and give the \( \Delta H \) for each:
   a. \( \text{C}_2\text{H}_4(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \) \( \Delta H = +2220 \text{ kJ} \)
   b. \( 2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s) \) \( \Delta H = +819 \text{ kJ} \)
   c. \( \text{PCl}_3(g) + 67 \text{ kJ} \rightarrow \text{PCl}_3(g) + \text{Cl}_2(g) \) \( \Delta H = +67 \text{ kJ} \)

9.39 a. How many kilojoules are released when 125 g of \( \text{Cl}_2 \) reacts with silicon?
   \( \text{Si}(s) + 2\text{Cl}_2(g) \rightarrow \text{SiCl}_4(g) \) \( \Delta H = -657 \text{ kJ} \)
   b. How many kilojoules are absorbed when 278 g of \( \text{PCl}_3 \) reacts?
   \( \text{PCl}_3(g) \rightarrow \text{PCl}_3(g) + \text{Cl}_2(g) \) \( \Delta H = +67 \text{ kJ} \)

9.40 a. How many kilojoules are released when 75.0 g of methanol reacts?
   \( 2\text{CH}_3\text{O}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 4\text{H}_2\text{O}(l) \) \( \Delta H = -726 \text{ kJ} \)
   b. How many kilojoules are absorbed when 315 g of \( \text{Ca(OH)}_2 \) reacts?
   \( \text{Ca(OH)}_2(s) \rightarrow \text{CaO}(s) + \text{H}_2\text{O}(l) \) \( \Delta H = +65.3 \text{ kJ} \)

9.41 Calculate the enthalpy change for the reaction
   \( \text{N}_2(g) + 2\text{O}_2(g) \rightarrow 2\text{NO}_2(g) \) from the following:
   \( \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \) \( \Delta H = +180 \text{ kJ} \)
   \( 2\text{NO}_2(g) \rightarrow 2\text{NO}(g) + \text{O}_2(g) \) \( \Delta H = +112 \text{ kJ} \)

9.42 Calculate the enthalpy change for the reaction
   \( \text{Fe}_2\text{O}_3(s) + \text{CO}(g) \rightarrow 2\text{FeO}(s) + \text{CO}_2(g) \) from the following:
   \( \text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g) \) \( \Delta H = -23.4 \text{ kJ} \)
   \( \text{FeO}(s) + \text{CO}(g) \rightarrow \text{Fe}(s) + \text{CO}_2(g) \) \( \Delta H = -10.9 \text{ kJ} \)

9.43 Calculate the enthalpy change for the reaction
   \( \text{S}(s) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{SO}_3(g) \) from the following:
   \( \text{S}(s) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{SO}_3(g) \) \( \Delta H = -396 \text{ kJ} \)
   \( \text{SO}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{SO}_3(g) \) \( \Delta H = -90 \text{ kJ} \)

9.44 Calculate the enthalpy for the reaction
   \( 3\text{C}(s) + \text{O}_2(g) \rightarrow \text{C}_3\text{O}_2(g) \) from the following:
   \( 2\text{CO}(g) + \text{C}(s) \rightarrow \text{C}_3\text{O}_2(g) \) \( \Delta H = +127 \text{ kJ} \)
   \( \text{C}(s) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{CO}(g) \) \( \Delta H = -111 \text{ kJ} \)
One of the problems that Lance monitored was water pollution by insecticides that exceed government regulations. These insecticides are made by organic synthesis, in which smaller molecules are combined to form larger molecules, in a stepwise fashion.

9.45 In one step in the synthesis of the insecticide Sevin, naphthol reacts with phosgene as shown.

\[ \text{C}_{10}\text{H}_8\text{O} + \text{COCl}_2 \rightarrow \text{C}_{11}\text{H}_7\text{O}_2\text{Cl} + \text{HCl} \]

**Naphthol**  **Phosgene**

a. How many kilograms of \( \text{C}_{11}\text{H}_7\text{O}_2\text{Cl} \) form from \( 2.2 \times 10^2 \text{ kg} \) of naphthol?

b. If 100. g of naphthol and 100. g of phosgene react, what is the theoretical yield of \( \text{C}_{11}\text{H}_7\text{O}_2\text{Cl} \)?

c. If the actual yield of \( \text{C}_{11}\text{H}_7\text{O}_2\text{Cl} \) in part b is 115 g, what is the percent yield?

9.46 Another widely used insecticide is carbofuran (Furadan), an extremely toxic insecticide. In one step in the synthesis of carbofuran, the reaction shown is used.

\[ \text{C}_6\text{H}_6\text{O}_2 + \text{C}_4\text{H}_7\text{Cl} \rightarrow \text{C}_{10}\text{H}_{12}\text{O}_2 + \text{HCl} \]

a. How many grams of \( \text{C}_6\text{H}_6\text{O}_2 \) are needed to produce \( 3.8 \times 10^3 \text{ g} \) of \( \text{C}_{10}\text{H}_{12}\text{O}_2 \)?

b. If 67.0 g of \( \text{C}_6\text{H}_6\text{O}_2 \) and 51.0 g of \( \text{C}_4\text{H}_7\text{Cl} \) react, what is the theoretical yield of \( \text{C}_{10}\text{H}_{12}\text{O}_2 \)?

c. If the actual yield of \( \text{C}_{10}\text{H}_{12}\text{O}_2 \) in part b is 85.7 g, what is the percent yield?

---

### Concept Map

**Chemical Quantities in Reactions**

- **Use**
  - **Coefficients**
  - **Balanced Chemical Equation**
- **Use a**
  - **Mole-Mole Factors**
  - **Molar Mass**
- **Use**
  - **Heat of Reaction \( \Delta H \)**
  - **Grams of Reactants or Products**
- **That has**
  - **That indicates an**
    - **Exothermic Reaction**
    - **Endothermic Reaction**
    - **Heat Flows Out**
    - **Heat Flows In**
- **To calculate**
  - **Limiting Reactant**
  - **Theoretical Yield**
  - **Percent Yield**
  - **Hess’s Law**
CHAPTER REVIEW

9.1 Conservation of Mass

LEARNING GOAL Calculate the total mass of reactants and the total mass of products in a balanced chemical equation.

- In a balanced equation, the total mass of the reactants is equal to the total mass of the products.

9.2 Calculating Moles Using Mole–Mole Factors

LEARNING GOAL Use a mole–mole factor from a balanced chemical equation to calculate the number of moles of another substance in the reaction.

- The coefficients in an equation describing the relationship between the moles of any two components are used to write mole–mole factors.
- When the number of moles for one substance is known, a mole–mole factor is used to find the moles of a different substance in the reaction.

9.3 Mass Calculations for Reactions

LEARNING GOAL Given the mass in grams of a substance in a reaction, calculate the mass in grams of another substance in the reaction.

- In calculations using equations, the molar masses of the substances and their mole–mole factors are used to change the number of grams of one substance to the corresponding grams of a different substance.

9.4 Limiting Reactants

LEARNING GOAL Identify a limiting reactant when given the quantities of two reactants; calculate the amount of product formed from the limiting reactant.

- A limiting reactant is the reactant that produces the smaller amount of product while the other reactant is left over.
- When the masses of two reactants are given, the mass of a product is calculated from the limiting reactant.

9.5 Percent Yield

LEARNING GOAL Given the actual quantity of product, determine the percent yield for a reaction.

- The percent yield for a reaction indicates the percent of product actually produced during a reaction.
- The percent yield is calculated by dividing the actual yield in grams of a product by the theoretical yield in grams and multiplying by 100%.

9.6 Energy in Chemical Reactions

LEARNING GOAL Given the heat of reaction (enthalpy change), calculate the loss or gain of heat for an exothermic or endothermic reaction.

- In chemical reactions, the heat of reaction ($\Delta H$) is the energy difference between the products and the reactants.
- In an exothermic reaction, the energy of the products is lower than that of the reactants. Heat is released, and $\Delta H$ is negative.
- In an endothermic reaction, the energy of the products is higher than that of the reactants; heat is absorbed, and $\Delta H$ is positive.
- Hess’s law states that heat can be absorbed or released in a single step or in several steps.

KEY TERMS

actual yield The actual amount of product produced by a reaction.
endothermic reaction A reaction wherein the energy of the products is higher than that of the reactants.
exothermic reaction A reaction wherein the energy of the products is lower than that of the reactants.
heat of reaction The heat (symbol $\Delta H$) absorbed or released when a reaction takes place at constant pressure.
Hess’s law Heat can be absorbed or released in a single chemical reaction or in several steps.
law of conservation of mass In a chemical reaction, the total mass of the reactants is equal to the total mass of the products; matter is neither lost nor gained.

limiting reactant The reactant used up during a chemical reaction, which limits the amount of product that can form.
mole–mole factor A conversion factor that relates the number of moles of two compounds in an equation derived from their coefficients.
percent yield The ratio of the actual yield for a reaction to the theoretical yield possible for the reaction.
theoretical yield The maximum amount of product that a reaction can produce from a given amount of reactant.
CORE CHEMISTRY SKILLS

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

Using Mole–Mole Factors (9.2)

Consider the balanced chemical equation

\[ 4Na(s) + O_2(g) \rightarrow 2Na_2O(s) \]

- The coefficients in a balanced chemical equation represent the moles of reactants and the moles of products. Thus, 4 mol of Na react with 1 mol of \( O_2 \) to form 2 mol of \( Na_2O \).
- From the coefficients, mole–mole factors can be written for any two substances as follows:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Mole–mole Factor</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na and ( O_2 )</td>
<td>[ \frac{4 \text{ mol } Na}{1 \text{ mol } O_2} ] and [ \frac{1 \text{ mol } O_2}{4 \text{ mol } Na} ]</td>
</tr>
<tr>
<td>Na and ( Na_2O )</td>
<td>[ \frac{4 \text{ mol } Na}{2 \text{ mol } Na_2O} ] and [ \frac{2 \text{ mol } Na_2O}{4 \text{ mol } Na} ]</td>
</tr>
<tr>
<td>( O_2 ) and ( Na_2O )</td>
<td>[ \frac{1 \text{ mol } O_2}{2 \text{ mol } Na_2O} ] and [ \frac{2 \text{ mol } Na_2O}{1 \text{ mol } O_2} ]</td>
</tr>
</tbody>
</table>

- A mole–mole factor is used to convert the number of moles of one substance in the reaction to the number of moles of another substance in the reaction.

**Example:** How many moles of sodium are needed to produce 3.5 mol of sodium oxide?

**Answer:**

\[ 3.5 \text{ mol } Na_2O \times \frac{4 \text{ mol } Na}{2 \text{ mol } Na_2O} = 7.0 \text{ mol of Na} \]

Two SFs

Converting Grams to Grams (9.3)

- When we have the balanced chemical equation for a reaction, we can use the mass of substance A and then calculate the mass of substance B. The process is as follows:
  - Use the molar mass of A to convert the mass, in grams, of A to moles of A.
  - Use the mole–mole factor that converts moles of A to moles of B.
  - Use the molar mass of B to calculate the mass, in grams, of B.

Example: How many grams of \( O_2 \) are needed to completely react with 14.6 g of Na?

\[ 4Na(s) + O_2(g) \rightarrow 2Na_2O(s) \]

**Answer:**

\[ 14.6 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \times \frac{1 \text{ mol } O_2}{4 \text{ mol Na}} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 5.08 \text{ g of } O_2 \]

Three SFs

Calculating Quantity of Product from a Limiting Reactant (9.4)

Often in reactions, the reactants are not present in quantities that allow both reactants to be completely used up. Then one of the reactants, called the limiting reactant, determines the maximum amount of product that can form.

- To determine the limiting reactant, we calculate the amount of product that is possible from each reactant.
- The limiting reactant is the one that produces the smaller amount of product.

**Example:** If 12.5 g of S reacts with 17.2 g of \( O_2 \), what is the limiting reactant and the mass, in grams, of \( SO_3 \) produced?

\[ 2S(s) + 3O_2(g) \rightarrow 2SO_3(g) \]

**Answer:**

**Mass of \( SO_3 \) from \( S \):**

\[ 12.5 \text{ g } S \times \frac{1 \text{ mol } S}{32.07 \text{ g } S} \times \frac{2 \text{ mol } SO_3}{2 \text{ mol } S} = 80.07 \text{ g } SO_3 \]

Three SFs

**Mass of \( SO_3 \) from \( O_2 \):**

\[ 17.2 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{3 \text{ mol } SO_3}{2 \text{ mol } O_2} = 80.07 \text{ g } SO_3 \]

Three SFs

Limiting reactant = 28.7 g of \( SO_3 \) Smaller amount of \( SO_3 \)

Calculating Percent Yield (9.5)

- The theoretical yield for a reaction is the amount of product (100%) formed if all the reactants were converted to desired product.
- The actual yield for the reaction is the mass, in grams, of the product obtained at the end of the experiment. Because some product is usually lost, the actual yield is less than the theoretical yield.
- The percent yield is calculated from the actual yield divided by the theoretical yield and multiplied by 100%.

\[ \text{Percent yield } (\%) = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \]

**Example:** If 22.6 g of Al reacts completely with \( O_2 \), and 37.8 g of \( Al_2O_3 \) is obtained, what is the percent yield of \( Al_2O_3 \) for the reaction?

\[ 4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s) \]

**Answer:**

**Calculation of theoretical yield:**

\[ 22.6 \text{ g } Al \times \frac{1 \text{ mol } Al}{26.98 \text{ g } Al} \times \frac{2 \text{ mol } Al_2O_3}{4 \text{ mol } Al} \times \frac{101.96 \text{ g } Al_2O_3}{1 \text{ mol } Al_2O_3} = 42.7 \text{ g of } Al_2O_3 \] Theoretical yield

Three SFs

**Calculation of percent yield:**

\[ \frac{\text{actual yield (given)}}{\text{theoretical yield (calculated)}} \times 100\% \]

\[ = \frac{37.8 \text{ g } Al_2O_3}{42.7 \text{ g } Al_2O_3} \times 100\% = 88.5\% \]

Three SFs
Using the Heat of Reaction (9.6)

- The heat of reaction is the amount of heat, usually in kJ, that is absorbed or released during a reaction.
- The heat of reaction or enthalpy change, symbol \( \Delta H \), is the difference in the energy of the products and the reactants.
  \[ \Delta H = H_{\text{products}} - H_{\text{reactants}} \]
- In an exothermic reaction (\( \text{exo} \) means “out”), the energy of the products is lower than that of the reactants. This means that heat is released along with the products that form. Then the sign for the heat of reaction, \( \Delta H \), is negative.
- In an endothermic reaction (\( \text{endo} \) means “within”), the energy of the products is higher than that of the reactants. The heat is required to convert the reactants to products. Then the sign for the heat of reaction, \( \Delta H \), is positive.

**Example:** How many kilojoules are released when 3.50 g of \( \text{CH}_4 \) undergoes combustion?

\[
\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g) \quad \Delta H = -802 \text{ kJ}
\]

**Answer:**

\[
3.50 \text{ g } \text{CH}_4 \times \frac{1 \text{ mol } \text{CH}_4}{16.04 \text{ g } \text{CH}_4} \times \frac{-802 \text{ kJ}}{1 \text{ mol } \text{CH}_4} = -175 \text{ kJ}
\]

**UNDERSTANDING THE CONCEPTS**

_The chapter sections to review are shown in parentheses at the end of each question._

9.47 If red spheres represent oxygen atoms and blue spheres represent nitrogen atoms, and all the molecules are gases, (9.2, 9.4)

**Reactants**

**Products**

a. write a balanced equation for the reaction
b. identify the limiting reactant

9.48 If green spheres represent chlorine atoms, yellow-green spheres represent fluorine atoms, and white spheres represent hydrogen atoms, and all the molecules are gases, (9.2, 9.4)

**Reactants**

**Products**

a. write a balanced equation for the reaction
b. identify the limiting reactant

9.49 If blue spheres represent nitrogen atoms and white spheres represent hydrogen atoms, and all the molecules are gases, (9.2, 9.4)

**Reactants**

**Products**

a. write a balanced equation for the reaction
b. identify the diagram that shows the products
9.50 If purple spheres represent iodine atoms and white spheres represent hydrogen atoms, and all the molecules are gases, (9.2, 9.4)

a. write a balanced equation for the reaction

b. identify the diagram that shows the products

9.51 If blue spheres represent nitrogen atoms and purple spheres represent iodine atoms, and the reacting molecules are solid, and the products are gases, (9.2, 9.4, 9.5)

a. write a balanced equation for the reaction

b. from the diagram of the actual products that result, calculate the percent yield for the reaction

9.52 If green spheres represent chlorine atoms and red spheres represent oxygen atoms, and all the molecules are gases, (9.2, 9.4, 9.5)

a. write a balanced equation for the reaction

b. identify the limiting reactant

c. from the diagram of the actual products that result, calculate the percent yield for the reaction

9.53 Use the balanced chemical equation to complete the table: (9.1, 9.2)

\[ \text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g) \]

<table>
<thead>
<tr>
<th>(\text{Fe}_2\text{O}_3)</th>
<th>(\text{CO})</th>
<th>(\text{Fe})</th>
<th>(\text{CO}_2)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.5 mol</td>
<td>_____ mol</td>
<td>_____ mol</td>
<td>_____ mol</td>
</tr>
<tr>
<td>_____ mol</td>
<td>_____ mol</td>
<td>_____ mol</td>
<td>0.6 mol</td>
</tr>
</tbody>
</table>

9.54 Use the balanced chemical equation to complete the table: (9.1, 9.2)

\[ 4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g) \]

<table>
<thead>
<tr>
<th>(\text{NH}_3)</th>
<th>(\text{O}_2)</th>
<th>(\text{NO})</th>
<th>(\text{H}_2\text{O})</th>
</tr>
</thead>
<tbody>
<tr>
<td>_____ mol</td>
<td>1 mol</td>
<td>_____ mol</td>
<td>_____ mol</td>
</tr>
<tr>
<td>_____ mol</td>
<td>_____ mol</td>
<td>2.8 mol</td>
<td>_____ mol</td>
</tr>
</tbody>
</table>
9.55 When ammonia (NH₃) reacts with fluorine (F₂), the products are dinitrogen tetrafluoride and hydrogen fluoride. (9.2, 9.3)
\[ 2\text{NH}_3(g) + 5\text{F}_2(g) \rightarrow \text{N}_2\text{F}_4(g) + 6\text{HF}(g) \]

a. How many moles of each reactant are needed to produce 4.00 mol of HF?
b. How many grams of F₂ are required to react with 25.5 g of NH₃?
c. How many grams of N₂F₄ can be produced when 3.40 g of NH₃ reacts?

9.56 Diesel contains hydrocarbon (C₁₀H₂₂) that burns in oxygen (O₂) to give carbon dioxide and water. (9.2, 9.3)
\[ 2\text{C}_10\text{H}_{22}(aq) + 31\text{O}_2(g) \rightarrow 20\text{CO}_2(g) + 22\text{H}_2\text{O}(l) \]

a. How many moles of O₂ are needed to completely react with 1.0 mol of C₁₀H₂₂?
b. If a car produces 44 g of CO₂, how many grams of C₁₀H₂₂ are used in the reaction?
c. If you add 28.8 g of C₁₀H₂₂ to your fuel, how many moles of O₂ are used up in the reaction?

9.57 When hydrogen peroxide (H₂O₂) is used in rocket fuels, it produces water and oxygen (O₂). (9.2, 9.3)
\[ 2\text{H}_2\text{O}_2(l) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g) \]

a. How many moles of H₂O₂ are needed to produce 3.00 mol of H₂O?
b. How many grams of H₂O₂ are required to produce 36.5 g of O₂?
c. How many grams of H₂O can be produced when 12.2 g of H₂O₂ reacts?

9.58 Propane gas, C₃H₈, reacts with oxygen to produce carbon dioxide and water. (9.2, 9.3)
\[ \text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \]

a. How many moles of H₂O form when 5.00 mol of C₃H₈ completely reacts?
b. How many grams of CO₂ are produced from 18.5 g of oxygen gas?
c. How many grams of H₂O can be produced when 46.3 g of C₃H₈ reacts?

9.59 When 72.9 g of Mg and 30 g of N₂ react, what is the mass, in grams, of Mg₃N₂ that is produced? (9.2, 9.3, 9.4)
\[ 3\text{Mg}(s) + \text{N}_2(g) \rightarrow \text{Mg}_3\text{N}_2(s) \]

9.60 When 65.4 g of Zn and 18 g of O₂ react, what is the mass, in grams, of ZnO that is produced? (9.2, 9.3, 9.4)
\[ 2\text{Zn}(s) + \text{O}_2(g) \rightarrow 2\text{ZnO}(s) \]

9.61 Pentane gas, C₅H₁₂, reacts with oxygen to produce carbon dioxide and water. (9.2, 9.3, 9.4)
\[ \text{C}_5\text{H}_{12}(g) + 8\text{O}_2(g) \rightarrow 5\text{CO}_2(g) + 6\text{H}_2\text{O}(g) \]

a. How many moles of C₅H₁₂ must react to produce 4.00 mol of water?
b. How many grams of CO₂ are produced from 32.0 g of O₂?
c. How many grams of O₂ are formed if 44.5 g of C₅H₁₂ is mixed with 108 g of O₂?

9.62 When nitrogen dioxide (NO₂) from car exhaust combines with water in the air, it forms nitrogen oxide and nitric acid (HNO₃), which causes acid rain. (9.2, 9.3, 9.4)
\[ 3\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{NO}(g) + 2\text{HNO}_3(aq) \]

a. How many moles of NO₂ are needed to react with 0.250 mol of H₂O?
b. How many grams of HNO₃ are produced when 60.0 g of NO₂ completely reacts?
c. How many grams of HNO₃ can be produced if 225 g of NO₂ is mixed with 55.2 g of H₂O?

9.63 The gaseous hydrocarbon acetylene, C₂H₂, used in welders’ torches, burns according to the following equation: (9.2, 9.3, 9.4, 9.5)
\[ 2\text{C}_2\text{H}_2(g) + 5\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 2\text{H}_2\text{O}(g) \]

a. What is the theoretical yield, in grams, of CO₂, if 22.0 g of C₂H₂ completely reacts?
b. If the actual yield in part a is 64.0 g of CO₂, what is the percent yield of CO₂ for the reaction?

9.64 The equation for the decomposition of potassium chlorate is written as (9.2, 9.3, 9.4, 9.5)
\[ 2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g) \]

a. When 46.0 g of KClO₃ is completely decomposed, what is the theoretical yield, in grams, of O₂?
b. If the actual yield in part a is 12.1 g of O₂, what is the percent yield of O₂?

9.65 When 4.2 g of ethene (C₂H₄) reacts with hydrogen, 3.8 g of ethane is formed. What is the percentage yield of C₂H₆ for the reaction? (9.2, 9.3, 9.4, 9.5)
\[ \text{C}_2\text{H}_4(g) + \text{H}_2(g) \rightarrow \text{C}_2\text{H}_6(g) \]

9.66 When 47.2 g of iron(III) oxide, reacts with carbon, 19.8 g of iron is produced. What is the percentage yield of Fe for the reaction? (9.2, 9.3, 9.4, 9.5)
\[ 2\text{Fe}_2\text{O}_3(s) + 3\text{C}(s) \rightarrow 4\text{Fe}(s) + 3\text{CO}_2(g) \]

9.67 Nitrogen and hydrogen combine to form ammonia. (9.2, 9.3, 9.4, 9.5)
\[ \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \]

a. If 50.0 g of N₂ is mixed with 20.0 g of H₂, what is the theoretical yield, in grams, of NH₃?
b. If the reaction in part a has a percent yield of 62.0%, what is the actual yield, in grams, of ammonia?

9.68 Sodium and nitrogen combine to form sodium nitride. (9.2, 9.3, 9.4, 9.5)
\[ 6\text{Na}(s) + \text{N}_2(g) \rightarrow 2\text{Na}_3\text{N}(s) \]

a. If 80.0 g of Na is mixed with 20.0 g of nitrogen gas, what is the theoretical yield, in grams, of Na₃N?
b. If the reaction in part a has a percent yield of 75.0%, what is the actual yield, in grams, of Na₃N?

9.69 The equation for the reaction of nitrogen and oxygen to form nitrogen oxide is written as (9.2, 9.6)
\[ \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \quad \Delta H = +90.2 \text{ kJ} \]

a. How many kilojoules are required to form 3.00 g of NO?
b. What is the complete equation (including heat) for the decomposition of NO?
c. How many kilojoules are released when 5.00 g of NO decomposes to N₂ and O₂?

9.70 The equation for the reaction of iron and oxygen gas to form rust (Fe₂O₃) is written as (9.2, 9.6)
\[ 4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \quad \Delta H = -1.7 \times 10^3 \text{ kJ} \]

a. How many kilojoules are released when 2.00 g of Fe reacts?
b. How many grams of rust form when 150 kJ are released?
9.71 Each of the following is a reaction that occurs in the cells of the body. Identify which is exothermic and endothermic. (9.6)

a. Succinyl CoA + H2O \rightarrow\text{succinate} + \text{CoA} + 37\text{kJ}

b. GDP + P_i + 34\text{kJ} \rightarrow\text{GTP} + \text{H}_2\text{O}

9.72 Each of the following is a reaction that occurs in the cells of the body. Identify which is exothermic and endothermic. (9.6)

a. Phosphocreatine + H2O \rightarrow \text{creatine} + P_i + 42.7\text{kJ}

b. Fructose-6-phosphate + P_i + 16\text{kJ} \rightarrow \text{fructose-1,6-bisphosphate}

c. \text{If the reaction in part c produces} 145\text{g of CO}_2, \text{what is the percent yield of CO}_2 \text{for the reaction?}

9.73 Chromium and oxygen combine to form chromium(III) oxide. (9.2, 9.3, 9.4, 9.5)

\[4\text{Cr(s)} + 3\text{O}_2(g) \rightarrow 2\text{Cr}_2\text{O}_3(s)\]

a. How many moles of O2 react with 4.50 mol of Cr?

b. How many grams of Cr2O3 are produced when 24.8 g of Cr reacts?

c. When 26.0 g of Cr reacts with 8.00 g of O2, how many grams of Cr2O3 can form?

d. If 74.0 g of Cr and 62.0 g of O2 are mixed, and 87.3 g of Cr2O3 is actually obtained, what is the percent yield of Cr2O3 for the reaction?

9.74 Aluminum and chlorine combine to form aluminum chloride. (9.2, 9.3, 9.4, 9.5)

\[2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s)\]

a. How many moles of Cl2 are needed to react with 4.50 mol of Al?

b. How many grams of AlCl3 are produced when 50.2 g of Al reacts?

c. When 13.5 g of Al reacts with 8.00 g of Cl2, how many grams of AlCl3 can form?

d. If 45.0 g of Al and 62.0 g of Cl2 are mixed, and 66.5 g of AlCl3 is actually obtained, what is the percent yield of AlCl3 for the reaction?

9.75 The combustion of propyne, C3H4, releases heat when it burns according to the following equation: (9.2, 9.3, 9.4, 9.5)

\[\text{C}_3\text{H}_4(g) + 4\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 2\text{H}_2\text{O}(g)\]

a. How many moles of O2 are needed to react completely with 0.225 mol of C3H4?

b. How many grams of water are produced from the complete reaction of 64.0 g of O2?

c. How many grams of CO2 are produced from the complete reaction of 78.0 g of C3H4?

d. If the reaction in part e produces 186 g of CO2, what is the percent yield of CO2 for the reaction?

9.76 Butane gas, C4H10, burns according to the following equation: (9.2, 9.3, 9.4, 9.5)

\[2\text{C}_4\text{H}_10(g) + 13\text{O}_2(g) \rightarrow 8\text{CO}_2(g) + 10\text{H}_2\text{O}(g)\]

a. How many moles of H2O are produced from the complete reaction of 2.50 mol of C4H10?

b. How many grams of O2 are needed to react completely with 22.5 g of C4H10?

c. How many grams of CO2 are produced from the complete reaction of 55.0 g of C4H10?

9.77 Use the balanced chemical equation to complete the table: (9.1, 9.2, 9.3)

\[\text{Zn(s)} + 2\text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)\]

<table>
<thead>
<tr>
<th>Zn</th>
<th>HCl</th>
<th>ZnCl2</th>
<th>H2</th>
</tr>
</thead>
<tbody>
<tr>
<td>___ g</td>
<td>___ g</td>
<td>50.5 g</td>
<td>___ g</td>
</tr>
<tr>
<td>___ g</td>
<td>___ g</td>
<td>___ g</td>
<td>0.4 g</td>
</tr>
</tbody>
</table>

9.78 Use the balanced chemical equation to complete the table: (9.1, 9.2, 9.3)

\[2\text{MnO}_2(s) + \text{CO}(g) \rightarrow \text{Mn}_2\text{O}_3(s) + \text{CO}_2(g)\]

<table>
<thead>
<tr>
<th>MnO2</th>
<th>CO</th>
<th>Mn2O3</th>
<th>CO2</th>
</tr>
</thead>
<tbody>
<tr>
<td>___ g</td>
<td>___ g</td>
<td>157.9 g</td>
<td>___ g</td>
</tr>
<tr>
<td>___ g</td>
<td>42.0 g</td>
<td>___ g</td>
<td>___ g</td>
</tr>
</tbody>
</table>

9.79 Sulfur reacts with carbon to form carbon disulfide. (9.2, 9.6)

\[\text{C}(s) + 2\text{S}(s) \rightarrow \text{CS}_2(g)\] \[\Delta H = +92\text{kJ}\]

a. Is the reaction endothermic or exothermic?

b. How many kilojoules are required when 1.5 mol of S reacts?

c. How many kilojoules are required when 100 g of CS2 is formed?

9.80 When hydrogen peroxide (H2O2) is used in rocket fuels, it produces water, oxygen, and heat. (9.2, 9.6)

\[2\text{H}_2\text{O}_2(l) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g)\] \[\Delta H = -196\text{kJ}\]

a. Is the reaction endothermic or exothermic?

b. How many kilojoules are released when 2.50 mol of H2O2 reacts?

c. How many kilojoules are released when 275 g of O2 is produced?

9.81 Calculate the enthalpy change for the reaction

\[\text{NH}_4\text{Cl}(s) \rightarrow \text{NH}_3(g) + \text{HCl}(g)\]

from the following equations: (9.6)

\[\frac{1}{2}\text{H}_2(g) + \frac{1}{2}\text{Cl}_2(g) \rightarrow \text{HCl}(g)\] \[\Delta H = -92\text{kJ}\]

\[\text{N}_2(g) + 4\text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2\text{NH}_4\text{Cl}(s)\] \[\Delta H = -631\text{kJ}\]

\[\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)\] \[\Delta H = -296\text{kJ}\]

9.82 Calculate the enthalpy change for the reaction

\[\text{Mg}(s) + \text{N}_2(g) + 3\text{O}_2(g) \rightarrow \text{Mg(NO)}_3\text{(s)}\]

from the following equations: (9.6)

\[8\text{Mg}(s) + \text{Mg(NO)}_3\text{(s)} \rightarrow 3\text{Mg}_3\text{N}_2\text{(s)} + 6\text{MgO}(s)\] \[\Delta H = -3281\text{kJ}\]

\[\text{Mg}_3\text{N}_2\text{(s)} \rightarrow 3\text{Mg}(s) + \text{N}_2(g)\] \[\Delta H = +461\text{kJ}\]

\[2\text{MgO}(s) \rightarrow 2\text{Mg}(s) + \text{O}_2(g)\] \[\Delta H = +1204\text{kJ}\]
9.33 In exothermic reactions, the energy of the products is lower than that of the reactants.
9.35 a. exothermic b. endothermic c. exothermic
9.37 a. Heat is released, exothermic, $\Delta H = -802$ kJ
b. Heat is absorbed, endothermic, $\Delta H = +65.3$ kJ c. Heat is released, exothermic, $\Delta H = -850$ kJ
9.39 a. 579 kJ b. 89 kJ
9.41 +68 kJ
9.43 -306 kJ
9.45 a. $3.2 \times 10^2$ kg b. 143 g c. 80.4%
9.47 a. $2\text{NO(g)} + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$
b. NO is the limiting reactant.
9.49 a. $\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)$
b. A
9.51 a. $2\text{NH}_3(g) \rightarrow \text{N}_2(g) + 3\text{H}_2(g)$ b. 67%

### Answers to Selected Questions and Problems

9.1 a. 160.14 g of reactants = 160.14 g of products b. 283.88 g of reactants = 283.88 g of products

9.3 a. $\frac{2 \text{ mol SO}_2}{1 \text{ mol O}_2} \text{ and } \frac{1 \text{ mol O}_2}{2 \text{ mol SO}_2}$
   $\frac{2 \text{ mol SO}_2}{2 \text{ mol SO}_3} \text{ and } \frac{2 \text{ mol SO}_3}{1 \text{ mol O}_2}$
   $\frac{2 \text{ mol SO}_3}{2 \text{ mol SO}_3} \text{ and } \frac{1 \text{ mol O}_2}{2 \text{ mol SO}_3}$

b. $\frac{4 \text{ mol P}}{5 \text{ mol O}_2} \text{ and } \frac{5 \text{ mol O}_2}{4 \text{ mol P}}$
   $\frac{4 \text{ mol P}}{2 \text{ mol P}_2\text{O}_5} \text{ and } \frac{2 \text{ mol P}_2\text{O}_5}{4 \text{ mol P}}$
   $\frac{5 \text{ mol O}_2}{2 \text{ mol P}_2\text{O}_5} \text{ and } \frac{2 \text{ mol P}_2\text{O}_5}{5 \text{ mol O}_2}$

9.5 a. mol$\text{SO}_2$ × $\frac{2 \text{ mol SO}_3}{2 \text{ mol SO}_2}$ = mol of SO3
   b. mol$\text{P}$ × $\frac{5 \text{ mol O}_2}{4 \text{ mol P}}$ = mol of O2

9.7 a. 1.3 mol of O2 b. 10. mol of H2 c. 5.0 mol of H2O
9.9 a. 1.25 mol of C b. 0.96 mol of CO c. 1.0 mol of SO2 d. 0.50 mol of CS2
9.11 a. 77.5 g of Na2O b. 6.26 g of O2 c. 19.4 g of O2
9.13 a. 19.2 g of O2 b. 3.79 g of N2 c. 54.0 g of H2O
9.15 a. 3.66 g of H2O b. 26.3 g of NO c. 7.53 g of HNO3
9.17 a. $2\text{PbS(s)} + 3\text{O}_2(g) \rightarrow 2\text{PbO(s)} + 2\text{SO}_2(g)$
b. 6.00 g of O2 c. 17.4 g of SO2
   d. 137 g of PbS

9.19 a. Eight taxis can be used to pick up passengers. b. Seven taxis can be driven.
9.21 a. 5.0 mol of H2 b. 4.0 mol of H2 c. 3.0 mol of N2
9.23 a. C12 is the limiting reactant, which would produce 25.1 g of AlCl3.
   b. O2 is the limiting reactant, which would produce 13.5 g of H2O.
   c. O2 is the limiting reactant, which would produce 26.7 g of SO2.
9.25 a. 31.2 g of SO3 b. 34.6 g of Fe3O4 c. 35.0 g of CO2
9.27 a. 71.0% b. 63.2%
9.29 70.9 g of Al2O3
9.31 60.5%

9.93 AlCl3.
Bonding and Properties of Solids and Liquids

BILL HAS BEEN DIAGNOSED with basal cell carcinoma, the most common form of skin cancer. He has an appointment to undergo Mohs surgery, a specialized procedure to remove the cancerous growth found on his shoulder. The surgeon begins the process by removing the abnormal growth, in addition to a thin layer of surrounding (margin) tissue, which he sends to Lisa, a histologist. Lisa prepares the tissue sample to be viewed by a pathologist. Tissue preparation requires Lisa to cut the tissue into a very thin section (normally about 0.001 cm), which is then mounted onto a microscope slide.

Lisa then treats the tissue with a dye to stain the cells, as this enables the pathologist to view any abnormal cells more easily. The pathologist examines the tissue sample and reports back to the surgeon that no abnormal cells were present in the margin tissue and that Bill’s tumor has been completely removed. No further tissue removal is necessary.

CAREER

Histologist

Histologists study the microscopic make-up of tissues, cells, and bodily fluids in order to detect and identify the presence of a specific disease. They determine blood types and the concentrations of drugs and other substances in the blood. Histologists also help establish a rationale as to why a patient may not be responding to his or her treatment. Sample preparation is a critical component of a histologist’s job, as they prepare tissue samples from humans, animals, and plants. The tissue samples are cut into extremely thin sections, which are then mounted and stained using various chemical dyes. The dyes provide contrast for the cells to be viewed and help highlight any abnormalities that may exist. Utilization of various dyes requires the histologist to be familiar with solution preparation and the handling of potentially hazardous chemicals.
10.1 Lewis Structures for Molecules and Polyatomic Ions

**LEARNING GOAL** Draw the Lewis structures for molecular compounds or polyatomic ions with single and multiple bonds.

Now we can investigate more complex chemical bonds and how they contribute to the structure of a molecule or a polyatomic ion. Lewis structures use Lewis symbols to diagram the sharing of valence electrons in molecules and polyatomic ions. The presence of multiple bonds can be identified, and resonance structures can be drawn, if needed.

**Lewis Symbols**

A **Lewis symbol** is a convenient way to represent the valence electrons, which are shown as dots placed on the sides, top, or bottom of the symbol for the element. Lewis symbols for selected elements are given in **TABLE 10.1**.

---

**TABLE 10.1** Lewis Symbols for Selected Elements in Periods 1 to 4

<table>
<thead>
<tr>
<th>Number of Valence Electrons</th>
<th>1A (1)</th>
<th>2A (2)</th>
<th>3A (13)</th>
<th>4A (14)</th>
<th>5A (15)</th>
<th>6A (16)</th>
<th>7A (17)</th>
<th>8A (18)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>He⁺⁺</td>
</tr>
<tr>
<td>Li</td>
<td></td>
<td>Be⁺</td>
<td>B⁺⁺</td>
<td>C⁺⁺</td>
<td>N⁺⁺</td>
<td>O⁺⁺</td>
<td>F⁺⁺</td>
<td>Ne⁺⁺</td>
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<tr>
<td>Na</td>
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<td>Al⁺</td>
<td>Si⁺⁺</td>
<td>P⁺⁺</td>
<td>S⁺⁺</td>
<td>Cl⁺⁺</td>
<td>Ar⁺⁺</td>
<td></td>
</tr>
<tr>
<td>K</td>
<td>Ca⁺⁺</td>
<td>Sr⁺⁺</td>
<td>Ca⁺⁺</td>
<td>Ge⁺⁺</td>
<td>As⁺⁺</td>
<td>Se⁺⁺</td>
<td>Br⁺⁺</td>
<td>Kr⁺⁺</td>
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</tbody>
</table>

*Helium (He) is stable with two valence electrons.
Lewis Structure for the Hydrogen Molecule

The simplest molecule is hydrogen, $\text{H}_2$. When two H atoms are far apart, there is no attraction between them. As the H atoms move closer, the positive charge of each nucleus attracts the electron of the other atom. This attraction, which is greater than the repulsion between the valence electrons, pulls the H atoms closer until they share a pair of valence electrons (see FIGURE 10.1). The result is called a covalent bond, in which the shared electrons give the stable electron configuration of He to each of the H atoms. When the H atoms form $\text{H}_2$, they are more stable than two individual H atoms.

![Lewis structure for hydrogen molecule](image)

**FIGURE 10.1** A covalent bond forms as H atoms move close together to share electrons.

What determines the attraction between two H atoms?

Lewis Structures for Molecular Compounds

A molecule is represented by a Lewis structure in which the valence electrons of all the atoms are arranged to give octets except for hydrogen, which has two electrons. The shared electrons, or bonding pairs, are shown as two dots or a single line between atoms. The nonbonding pairs of electrons, or lone pairs, are placed on the outside. For example, a fluorine molecule, $\text{F}_2$, consists of two fluorine atoms, which are in Group 7A (17), each with seven valence electrons. In the Lewis structure for the $\text{F}_2$ molecule, each F atom achieves an octet by sharing its unpaired valence electron.

![Lewis structures for molecular compounds](image)

Hydrogen ($\text{H}_2$) and fluorine ($\text{F}_2$) are examples of nonmetal elements whose natural state is diatomic; that is, they contain two like atoms. The elements that exist as diatomic molecules are listed in **TABLE 10.2**.

**TABLE 10.2** Elements That Exist as Diatomic Molecules

<table>
<thead>
<tr>
<th>Diatomic Molecule</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{H}_2$</td>
<td>Hydrogen</td>
</tr>
<tr>
<td>$\text{N}_2$</td>
<td>Nitrogen</td>
</tr>
<tr>
<td>$\text{O}_2$</td>
<td>Oxygen</td>
</tr>
<tr>
<td>$\text{F}_2$</td>
<td>Fluorine</td>
</tr>
<tr>
<td>$\text{Cl}_2$</td>
<td>Chlorine</td>
</tr>
<tr>
<td>$\text{Br}_2$</td>
<td>Bromine</td>
</tr>
<tr>
<td>$\text{I}_2$</td>
<td>Iodine</td>
</tr>
</tbody>
</table>

Drawing Lewis Structures

The number of electrons that a nonmetal atom shares and the number of covalent bonds it forms are usually equal to the number of electrons it needs to achieve a stable electron configuration.
To draw the Lewis structure for CH₄, we first draw the Lewis symbols for carbon and hydrogen.

\[
\cdot \ddot{\text{C}} \cdot \quad \cdot \text{H}
\]

Then we determine the number of valence electrons needed for carbon and hydrogen. When a carbon atom shares its four electrons with four hydrogen atoms, carbon obtains an octet and each hydrogen atom is complete with two shared electrons. The Lewis structure is drawn with the carbon atom as the central atom, with the hydrogen atoms on each of the sides. The bonding pairs of electrons, which are single covalent bonds, may also be shown as single lines between the carbon atom and each of the hydrogen atoms.

**TABLE 10.3** gives examples of Lewis structures and molecular models for some molecules.

When we draw a Lewis structure for a molecule or polyatomic ion, we show the sequence of atoms, the bonding pairs of electrons shared between atoms, and the nonbonding or *lone pairs* of electrons. From the formula, we identify the central atom, which is the element that has the fewer atoms. Then, the central atom is bonded to the other atoms, as shown in Sample Problem 10.1.

### TABLE 10.3 Lewis Structures for Some Molecular Compounds

<table>
<thead>
<tr>
<th>Compound</th>
<th>Lewis Structures</th>
<th>Molecular Models</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₄</td>
<td>H&gt;C&lt;</td>
<td>Methane molecule</td>
</tr>
<tr>
<td>NH₃</td>
<td>H:N&lt;</td>
<td>Ammonia molecule</td>
</tr>
<tr>
<td>H₂O</td>
<td>H=O&lt;</td>
<td>Water molecule</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 10.1 Drawing Lewis Structures**

Draw the Lewis structure for PCl₃, phosphorus trichloride, used commercially to prepare insecticides and flame retardants.

**TRY IT FIRST**

#### SOLUTION

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>PCl₃</td>
<td>Lewis structure</td>
<td>total valence electrons</td>
</tr>
</tbody>
</table>

**STEP 1** Determine the arrangement of atoms. In PCl₃, the central atom is P because there is only one P atom.

P  Cl  Cl

**STEP 2** Determine the total number of valence electrons. We use the group number to determine the number of valence electrons for each of the atoms in the molecule.

<table>
<thead>
<tr>
<th>Atom</th>
<th>Group</th>
<th>Valence Electrons</th>
<th>Total</th>
</tr>
</thead>
<tbody>
<tr>
<td>P</td>
<td>5A (15)</td>
<td>5 e⁻</td>
<td>5 e⁻</td>
</tr>
<tr>
<td>Cl</td>
<td>7A (17)</td>
<td>7 e⁻</td>
<td>21 e⁻</td>
</tr>
</tbody>
</table>

Total valence electrons for PCl₃ = 26 e⁻
STEP 3 Attach each bonded atom to the central atom with a pair of electrons.
Each bonding pair can also be represented by a bond line.

\[
\begin{align*}
\text{Cl} & \text{P} \text{Cl} \quad \text{or} \\
\text{Cl} & \text{P} \quad \text{Cl}
\end{align*}
\]

Six electrons \(3 \times 2e^-\) are used to bond the central P atom to three Cl atoms. Twenty valence electrons are left.

26 valence \(e^-\) – 6 bonding \(e^-\) = 20 \(e^-\) remaining

STEP 4 Place the remaining electrons using single bonds to complete octets.

We use the remaining 20 electrons as lone pairs, which are placed around the outer Cl atoms and on the P atom, such that all the atoms have octets.

\[
\begin{align*}
\text{:Cl} & \text{P} \text{:Cl} \quad \text{or} \\
\text{:Cl} & \text{P} \quad \text{:Cl}
\end{align*}
\]

STUDY CHECK 10.1
Draw the Lewis structure for \(\text{Cl}_2\text{O}\).

ANSWER

\[
\begin{align*}
\text{:Cl} & \text{:O} & \text{:Cl} \quad \text{or} \\
\text{:Cl} & \text{O} & \text{:Cl}
\end{align*}
\]

SAMPLE PROBLEM 10.2 Drawing Lewis Structures for Polyatomic Ions

Sodium chlorite, \(\text{NaClO}_2\), is a component of mouthwashes, toothpastes, and contact lens cleaning solutions. Draw the Lewis structure for the chlorite ion, \(\text{ClO}_2^-\).

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>(\text{ClO}_2^-)</td>
<td>Lewis structure</td>
<td>total valence electrons</td>
<td></td>
</tr>
</tbody>
</table>

STEP 1 Determine the arrangement of atoms. For the polyatomic ion \(\text{ClO}_2^-\), the central atom is Cl because there is only one Cl atom. For a polyatomic ion, the atoms and electrons are placed in brackets, and the charge is written outside to the upper right.

\[
[\text{O} \quad \text{Cl} \quad \text{O}]^-
\]

STEP 2 Determine the total number of valence electrons. We use the group numbers to determine the number of valence electrons for each of the atoms in the ion. Because the ion has a negative charge, one more electron is added to the valence electrons.

<table>
<thead>
<tr>
<th>Element</th>
<th>Group</th>
<th>Atoms</th>
<th>Valence Electrons (=)</th>
<th>Total</th>
</tr>
</thead>
<tbody>
<tr>
<td>O (6\text{A}\ (16))</td>
<td>2 O</td>
<td>(6e^-)</td>
<td>(12e^-)</td>
<td></td>
</tr>
<tr>
<td>Cl (7\text{A}\ (17))</td>
<td>1 Cl</td>
<td>(7e^-)</td>
<td>(7e^-)</td>
<td></td>
</tr>
<tr>
<td>Ionic charge (negative) add</td>
<td>1 (e^-)</td>
<td>(1e^-)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Total valence electrons for \(\text{ClO}_2^-\) \(= \) \(20e^-\)

STEP 3 Attach each bonded atom to the central atom with a pair of electrons. Each bonding pair can also be represented by a line, which indicates a single bond.

\[
\begin{align*}
\text{O} & \text{Cl} \text{O}^- \quad \text{or} \\
\text{O} & \text{||Cl} \text{|| O}^-
\end{align*}
\]

Four electrons are used to bond the O atoms to the central Cl atom, which leaves 16 valence electrons.
**STEP 4** Place the remaining electrons using single bonds to complete octets. We use the remaining 16 electrons as lone pairs; 12 electrons are drawn as lone pairs to complete the octets of the O atoms.

\[
\begin{align*}
[\cdot\cdot\cdot\text{Cl}::\cdot\cdot\cdot] & \quad [\cdot\cdot\cdot\text{O}::\cdot\cdot\cdot] \\
\text{The remaining four electrons are placed as two lone pairs on the central Cl atom.}
\end{align*}
\]

\[
\begin{align*}
[\cdot\cdot\cdot\text{Cl}::\cdot\cdot\cdot] & \quad [\cdot\cdot\cdot\text{Cl}::\cdot\cdot\cdot] \\
\end{align*}
\]

**STUDY CHECK 10.2**

Draw the Lewis structure for the polyatomic ion \(\text{NH}_2^–\).

**ANSWER**

\[
\begin{align*}
[H::\text{N}::H] & \quad [H::\text{N}::H] \\
\end{align*}
\]

**Double and Triple Bonds**

Up to now, we have looked at bonding in molecules having only single bonds. In many molecular compounds, atoms share two or three pairs of electrons to complete their octets.

Double and triple bonds form when the number of valence electrons is not enough to complete the octets of all the atoms in the molecule. Then one or more lone pairs of electrons from the atoms attached to the central atom are shared with the central atom. A **double bond** occurs when two pairs of electrons are shared; in a **triple bond**, three pairs of electrons are shared. Atoms of carbon, oxygen, nitrogen, and sulfur are most likely to form multiple bonds.

Atoms of hydrogen and the halogens do not form double or triple bonds. The process of drawing a Lewis structure with multiple bonds is shown in Sample Problem 10.3.

**SAMPLE PROBLEM 10.3** **Drawing Lewis Structures with Multiple Bonds**

Draw the Lewis structure for carbon dioxide, \(\text{CO}_2\), in which the central atom is C.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>**CO}_2</td>
<td>Lewis structure</td>
<td>total valence electrons</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Determine the arrangement of atoms. \(\text{O} \quad \text{C} \quad \text{O}\)

**STEP 2** Determine the total number of valence electrons. We use the group number to determine the number of valence electrons for each of the atoms in the molecule.

<table>
<thead>
<tr>
<th>Element</th>
<th>Group</th>
<th>Atoms</th>
<th>Valence Electrons</th>
<th>(=)</th>
<th>Total</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>4A (14)</td>
<td>1 C</td>
<td>(4 \ e^-)</td>
<td>(=)</td>
<td>(4 \ e^-)</td>
</tr>
<tr>
<td>O</td>
<td>6A (16)</td>
<td>2 O</td>
<td>(6 \ e^-)</td>
<td>(=)</td>
<td>(12 \ e^-)</td>
</tr>
</tbody>
</table>

Total valence electrons for \(\text{CO}_2\) \(=\) \(16 \ e^-\)
**STEP 3** Attach each bonded atom to the central atom with a pair of electrons.

\[
\text{O:C:O} \quad \text{or} \quad \text{O} \rightarrow \text{C} \rightarrow \text{O}
\]

We use four valence electrons to attach the central C atom to two O atoms.

**STEP 4** Place the remaining electrons using single or multiple bonds to complete octets.

The 12 remaining electrons are placed as six lone pairs of electrons on the outside O atoms. However, this does not complete the octet of the C atom.

\[
\text{O} \quad \text{or} \quad \text{O} \rightarrow \text{C} \rightarrow \text{O}
\]

To obtain an octet, the C atom must share lone pairs of electrons from each of the O atoms. When two bonding pairs occur between atoms, it is known as a double bond.

 Exceptions to the Octet Rule

Although the octet rule is useful for bonding in many compounds, there are exceptions. We have already seen that a hydrogen (H\(_2\)) molecule requires just two electrons or a single bond. Usually the nonmetals form octets. However, in BCl\(_3\), the B atom has only three valence electrons to share. Boron compounds typically have six valence electrons on the central B atoms and form just three bonds. Although we will generally see compounds of P, S, Cl, Br, and I with octets, they can form molecules in which they share more of their valence electrons. This expands their valence electrons to 10, 12, or even 14 electrons. For example, we have seen that the P atom in PCl\(_3\) has an octet, but in PCl\(_5\), the P atom has five bonds with 10 valence electrons. In H\(_2\)S, the S atom has an octet, but in SF\(_6\), there are six bonds to sulfur with 12 valence electrons.

**QUESTIONS AND PROBLEMS**

10.1 Lewis Structures for Molecules and Polyatomic Ions

**LEARNING GOAL** Draw the Lewis structures for molecular compounds or polyatomic ions with single and multiple bonds.

10.1 Determine the total number of valence electrons for each of the following:

| a. H\(_2\)S | b. I\(_2\) | c. CCl\(_4\) | d. OH\(^-\) |

10.2 Determine the total number of valence electrons for each of the following:

| a. SBr\(_2\) | b. NBr\(_3\) | c. CH\(_3\)OH | d. NH\(_4\)\(^+\) |

10.3 Draw the Lewis structure for each of the following molecules or polyatomic ions:

| a. HF | b. SF\(_2\) | c. NBr\(_3\) | d. BH\(_4\)\(^-\) |
| e. CH\(_3\)OH (methyl alcohol) | HCOH |
| f. N\(_2\)H\(_4\) (hydrazine) | HNNH |
10.4 Draw the Lewis structure for each of the following molecules or polyatomic ions:
   a. \( \text{H}_2\text{O} \)  
   b. \( \text{CCl}_4 \)  
   c. \( \text{H}_3\text{O}^+ \)  
   d. \( \text{SiF}_4 \)  
   e. \( \text{CF}_2\text{Cl}_2 \)  
   f. \( \text{C}_2\text{H}_6 \)  

10.5 When is it necessary to draw a multiple bond in a Lewis structure?

10.6 If the available number of valence electrons for a molecule or polyatomic ion does not complete all of the octets in a Lewis structure, what should you do?

10.7 Draw the Lewis structure for each of the following molecules or ions:
   a. \( \text{CO} \)  
   b. \( \text{CN}^- \)  
   c. \( \text{H}_2\text{CO} \) (\( \text{C} \) is the central atom)

10.8 Draw the Lewis structure for each of the following molecules or ions:
   a. \( \text{HCCH} \) (acetylene)  
   b. \( \text{CS}_2 \)  
   c. \( \text{NO}^+ \)

---

### 10.2 Resonance Structures

**LEARNING GOAL.** Draw Lewis structures for molecules or polyatomic ions that have two or more resonance structures.

When a molecule or polyatomic ion contains multiple bonds, it may be possible to draw more than one Lewis structure. We can see how this happens when we draw the Lewis structure for ozone, \( \text{O}_3 \), a component in the stratosphere that protects us from the ultraviolet rays of the Sun.

To draw the Lewis structure for \( \text{O}_3 \), we determine the number of valence electrons for an O atom, and then the total number of valence electrons for \( \text{O}_3 \). Because O is in Group 6A (16), it has six valence electrons. Therefore, the compound \( \text{O}_3 \) would have a total of 18 valence electrons.

<table>
<thead>
<tr>
<th>Element</th>
<th>Group</th>
<th>Atoms</th>
<th>Valence Electrons</th>
<th>=</th>
<th>Total</th>
</tr>
</thead>
<tbody>
<tr>
<td>\text{O}</td>
<td>6A (16)</td>
<td>3 O</td>
<td>( 6^- )</td>
<td>=</td>
<td>18 ( e^- )</td>
</tr>
</tbody>
</table>

For the Lewis structure for \( \text{O}_3 \), we place three O atoms in a row. Using four of the valence electrons, we draw a bonding pair between each of the O atoms on the ends and the central O atom. Two bonding pairs require four valence electrons.

\[
\text{O} \equiv \text{O} \equiv \text{O}
\]

The remaining valence electrons (14) are placed as lone pairs of electrons around the O atoms on both ends of the Lewis structure, and one lone pair goes on the central O atom.

\[
\overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}}
\]

To complete an octet for the central O atom, one lone pair of electrons from an end O atom needs to be shared. But which lone pair should be used? One possibility is to form a double bond between the central O atom and the O on the left, and the other possibility is to form a double bond between the central O atom and the O on the right.

\[
\overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}} \quad \text{or} \quad \overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}}
\]

Thus it is possible to draw two or more Lewis structures for a molecule such as \( \text{O}_3 \) or for a polyatomic ion. When this happens, all the Lewis structures are called resonance structures, and their relationship is shown by drawing a double-headed arrow between them.

\[
\overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}} \quad \leftrightarrow \quad \overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}} \equiv \overset{\cdot}{\text{O}}
\]

Resonance structures

Experiments show that the actual bond lengths in ozone are equivalent to a molecule with a “one-and-a-half” bond between the central O atom and each outside O atom. In an actual ozone molecule, the electrons are spread equally over all the O atoms. When we draw resonance structures for molecules or polyatomic ions, the true structure is really an average of those structures.
SAMPLE PROBLEM 10.4 Drawing Resonance Structures

Sulfur dioxide is produced by volcanic activity and the burning of sulfur-containing coal. Once in the atmosphere, the \( \text{SO}_2 \) is converted to \( \text{SO}_3 \), which combines with water, forming sulfuric acid, \( \text{H}_2\text{SO}_4 \), a component of acid rain. Draw two resonance structures for sulfur dioxide.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>( \text{SO}_2 )</td>
<td>resonance structures</td>
<td>total valence electrons, double bonds</td>
</tr>
</tbody>
</table>

**STEP 1** Determine the arrangement of atoms. In \( \text{SO}_2 \), the S atom is the central atom because there is only one S atom.

\[
\begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array}
\]

**STEP 2** Determine the total number of valence electrons. We use the group number to determine the number of valence electrons for each of the atoms in the molecule.

<table>
<thead>
<tr>
<th>Element</th>
<th>Group</th>
<th>Atoms</th>
<th>Valence Electrons</th>
<th>( \equiv )</th>
<th>Total</th>
</tr>
</thead>
<tbody>
<tr>
<td>S</td>
<td>6A (16)</td>
<td>1 S</td>
<td>( 6 \text{e}^- )</td>
<td>( = )</td>
<td>( 6 \text{e}^- )</td>
</tr>
<tr>
<td>O</td>
<td>6A (16)</td>
<td>2 O</td>
<td>( 6 \text{e}^- )</td>
<td>( = )</td>
<td>( 12 \text{e}^- )</td>
</tr>
</tbody>
</table>

\[
\text{Total valence electrons for } \text{SO}_2 = 18 \text{e}^-
\]

**STEP 3** Attach each bonded atom to the central atom with a pair of electrons.

\[
\begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array}
\]

We use four electrons to form single bonds between the central S atom and the O atoms.

**STEP 4** Place the remaining electrons using single or multiple bonds to complete octets. The remaining 14 electrons are drawn as lone pairs, which complete the octets for the O atoms but not the S atom.

\[
:\widetilde{\text{O}} \quad \text{S} \quad \text{O}: \\
\text{To complete the octet for S, one lone pair of electrons from one of the O atoms is shared to form a double bond. One possibility is to form a double bond between the central S atom and the O on the left, and the other possibility is to form a double bond between the central S atom and the O on the right.}
\]

\[
:\widetilde{\text{O}} \quad \text{S} \quad \text{O}: \quad \leftrightarrow \quad :\widetilde{\text{O}} \quad \text{S} = \text{O}: \\
\]

**STUDY CHECK 10.4**

Draw three resonance structures for \( \text{SO}_3 \).

**ANSWER**

\[
:\widetilde{\text{O}} \quad \text{S} \quad \text{O}: \quad \leftrightarrow \quad :\widetilde{\text{O}} \quad \text{S} \quad \text{O}: \quad \leftrightarrow \quad :\widetilde{\text{O}} \quad \text{S} = \text{O}: \\
\]

**ENGAGE**

Explain why \( \text{SO}_2 \) has resonance structures but \( \text{SCl}_2 \) does not.
TABLE 10.4 summarizes this method of drawing Lewis structures for several molecules and ions.

<table>
<thead>
<tr>
<th>Molecule or Polyatomic Ion</th>
<th>Total Valence Electrons</th>
<th>Form Single Bonds to Attach Atoms (electrons used)</th>
<th>Electrons Remaining</th>
<th>Completed Octets (or H)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl₂</td>
<td>2(7) = 14</td>
<td>Cl⁻ Cl⁺ (2 e⁻)</td>
<td>14 - 2 = 12</td>
<td>Cl⁻ Cl⁺</td>
</tr>
<tr>
<td>HCl</td>
<td>1 + 7 = 8</td>
<td>H⁻ Cl⁺ (2 e⁻)</td>
<td>8 - 2 = 6</td>
<td>H⁻ Cl⁺</td>
</tr>
<tr>
<td>H₂O</td>
<td>2(1) + 6 = 8</td>
<td>H⁻ O⁻ H⁺ (4 e⁻)</td>
<td>8 - 4 = 4</td>
<td>H⁻ O⁻ H⁺</td>
</tr>
<tr>
<td>PCl₃</td>
<td>5 + 3(7) = 26</td>
<td>Cl⁻ P⁺ Cl⁻ (6 e⁻)</td>
<td>26 - 6 = 20</td>
<td>Cl⁻ P⁺ Cl⁻</td>
</tr>
<tr>
<td>ClO₃⁻</td>
<td>7 + 3(6) + 1 = 26</td>
<td>O⁻ Cl⁺ O⁺</td>
<td>26 - 6 = 20</td>
<td>O⁻ Cl⁺ O⁺</td>
</tr>
<tr>
<td>NO₂⁻</td>
<td>5 + 2(6) + 1 = 18</td>
<td>O⁻ N⁺ O⁻ (4 e⁻)</td>
<td>18 - 4 = 14</td>
<td>O⁻ N⁺ O⁻</td>
</tr>
</tbody>
</table>

QUESTIONS AND PROBLEMS

10.2 Resonance Structures

LEARNING GOAL: Draw Lewis structures for molecules or polyatomic ions that have two or more resonance structures.

10.9 What is resonance?
10.10 When does a molecular compound have resonance?
10.11 Draw resonance structures for each of the following molecules or ions:
   a. ClNO₂ (N is the central atom)
   b. OCN⁻ (C is the central atom)

10.12 Draw resonance structures for each of the following molecules or ions:
   a. H₂CNO₂⁻
   b. N₂O (N N O)

10.3 Shapes of Molecules and Polyatomic Ions (VSEPR Theory)

LEARNING GOAL: Predict the three-dimensional structure of a molecule or a polyatomic ion.

Using the Lewis structures, we can predict the three-dimensional shapes of many molecules and polyatomic ions. The shape of a compound is important in our understanding of how molecules interact with enzymes or certain antibiotics or produce our sense of taste and smell.

The three-dimensional shape of a molecule is determined by drawing its Lewis structure and identifying the number of electron groups (one or more electron pairs) around the central atom. We count lone pairs of electrons, single, double, and triple bonds as one electron group. In the valence shell electron-pair repulsion (VSEPR) theory, the
electron groups are arranged as far apart as possible around the central atom to minimize the repulsion between their negative charges. Once we have counted the number of electron groups surrounding the central atom, we can determine its specific shape from the number of atoms bonded to the central atom.

**Central Atoms with Two Electron Groups**

In the Lewis structure for CO₂, there are two electron groups (two double bonds) attached to the central atom. According to VSEPR theory, minimal repulsion occurs when two electron groups are on opposite sides of the central C atom. This gives the CO₂ molecule a *linear* electron-group geometry and a shape that is *linear* with a bond angle of 180°.

\[ \text{\Linear}\ 
\]

**Central Atoms with Three Electron Groups**

In the Lewis structure for formaldehyde, H₂CO, the central atom C is attached to two H atoms by single bonds and to the O atom by a double bond. Minimal repulsion occurs when three electron groups are as far apart as possible around the central C atom, which gives 120° bond angles. This type of electron-group geometry is *trigonal planar* and gives a shape for H₂CO called *trigonal planar*.

\[ \text{\Trigonal}\ 
\]

In the Lewis structure for SO₂, there are also three electron groups around the central S atom: a single bond to an O atom, a double bond to another O atom, and a lone pair of electrons. As in H₂CO, three electron groups have minimal repulsion when they form trigonal planar electron-group geometry. However, in SO₂ one of the electron groups is a lone pair of electrons. Therefore, the shape of the SO₂ molecule is determined by the two O atoms bonded to the central S atom, which gives the SO₂ molecule a shape that is *bent* with a bond angle of 120°. When there are one or more lone pairs on the central atom, the shape has a different name than that of the electron-group geometry.

\[ \text{\Bent}\ 
\]
Central Atom with Four Electron Groups

In a molecule of methane, \( \text{CH}_4 \), the central C atom is bonded to four H atoms. From the Lewis structure, you may think that \( \text{CH}_4 \) is planar with 90° bond angles. However, the best geometry for minimal repulsion is tetrahedral, giving bond angles of 109°. When there are four atoms attached to four electron groups, the shape of the molecule is tetrahedral.

\[
\begin{align*}
\text{H} & \quad \text{H} & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{H} & & & &
\end{align*}
\]

Lewis structure

A way to represent the three-dimensional structure of methane is to use the wedge–dash notation. In this representation, the two bonds connecting carbon to hydrogen by solid lines are in the plane of the paper. The wedge represents a carbon-to-hydrogen bond coming out of the page toward us, whereas the dash represents a carbon-to-hydrogen bond going into the page away from us.

Now we can look at molecules that have four electron groups, of which one or more are lone pairs of electrons. Then the central atom is attached to only two or three atoms. For example, in the Lewis structure for ammonia, \( \text{NH}_3 \), four electron groups have a tetrahedral electron-group geometry. However, in \( \text{NH}_3 \) one of the electron groups is a lone pair of electrons. Therefore, the shape of \( \text{NH}_3 \) is determined by the three H atoms bonded to the central N atom. Therefore, the shape of the \( \text{NH}_3 \) molecule is trigonal pyramidal, with a bond angle of 109°. The wedge–dash notation can also represent this three-dimensional structure of ammonia with one N—H bond in the plane, one N—H bond coming toward us, and one N—H bond going away from us.

\[
\begin{align*}
\text{H} & \quad \text{N} & \quad \text{H} \\
\text{H} & & &
\end{align*}
\]

Trigonal pyramidal shape

In the Lewis structure for water, \( \text{H}_2\text{O} \), there are also four electron groups, which have minimal repulsion when the electron-group geometry is tetrahedral. However, in \( \text{H}_2\text{O} \), two of the electron groups are lone pairs of electrons. Because the shape of \( \text{H}_2\text{O} \) is determined by the two H atoms bonded to the central O atom, the \( \text{H}_2\text{O} \) molecule has a bent shape with a bond angle of 109°. TABLE 10.5 gives the shapes of molecules with two, three, or four bonded atoms.
**TABLE 10.5** Molecular Shapes for a Central Atom with Two, Three, and Four Bonded Atoms

<table>
<thead>
<tr>
<th>Electron Groups</th>
<th>Electron-Group Geometry</th>
<th>Bonded Atoms</th>
<th>Lone Pairs</th>
<th>Bond Angle*</th>
<th>Molecular Shape</th>
<th>Example</th>
<th>Three-Dimensional Model</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Linear</td>
<td>2</td>
<td>0</td>
<td>180°</td>
<td>Linear</td>
<td>CO₂</td>
<td><img src="image" alt="CO₂" /></td>
</tr>
<tr>
<td>3</td>
<td>Trigonal planar</td>
<td>3</td>
<td>0</td>
<td>120°</td>
<td>Trigonal planar</td>
<td>H₂CO</td>
<td><img src="image" alt="H₂CO" /></td>
</tr>
<tr>
<td>3</td>
<td>Trigonal planar</td>
<td>2</td>
<td>1</td>
<td>120°</td>
<td>Bent</td>
<td>SO₂</td>
<td><img src="image" alt="SO₂" /></td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral</td>
<td>4</td>
<td>0</td>
<td>109°</td>
<td>Tetrahedral</td>
<td>CH₄</td>
<td><img src="image" alt="CH₄" /></td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral</td>
<td>3</td>
<td>1</td>
<td>109°</td>
<td>Trigonal pyramidal</td>
<td>NH₃</td>
<td><img src="image" alt="NH₃" /></td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral</td>
<td>2</td>
<td>2</td>
<td>109°</td>
<td>Bent</td>
<td>H₂O</td>
<td><img src="image" alt="H₂O" /></td>
</tr>
</tbody>
</table>

*The bond angles in actual molecules may vary slightly.

---

**SAMPLE PROBLEM 10.5 Shapes of Molecules**

Use VSEPR theory to predict the shape of the molecule SiCl₄.

**TRY IT FIRST**

**SOLUTION**

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>SiCl₄</td>
<td>shape</td>
<td>Lewis structure, electron groups, bonded atoms</td>
</tr>
</tbody>
</table>

**STEP 1** Draw the Lewis structure.

<table>
<thead>
<tr>
<th>Name of Element</th>
<th>Silicon</th>
<th>Chlorine</th>
</tr>
</thead>
<tbody>
<tr>
<td>Symbol of Element</td>
<td>Si</td>
<td>Cl</td>
</tr>
<tr>
<td>Atoms of Element</td>
<td>1</td>
<td>4</td>
</tr>
<tr>
<td>Valence Electrons</td>
<td>4 e⁻</td>
<td>7 e⁻</td>
</tr>
<tr>
<td>Total Electrons</td>
<td>1(4 e⁻) + 4(7 e⁻) = 32 e⁻</td>
<td></td>
</tr>
</tbody>
</table>

Using 32 e⁻, we draw the bonds and lone pairs for the Lewis structure of SiCl₄.

![Lewis structure of SiCl₄]

**STEP 2** Arrange the electron groups around the central atom to minimize repulsion. To minimize repulsion, the four electron groups would have a tetrahedral geometry.

**Guide to Predicting Shape (VSEPR Theory)**

**STEP 1**

Draw the Lewis structure.

**STEP 2**

Arrange the electron groups around the central atom to minimize repulsion.

**STEP 3**

Use the atoms bonded to the central atom to determine the shape.
**STEP 3** Use the atoms bonded to the central atom to determine the shape.
Because the central Si atom is bonded to four atoms and no lone pairs of electrons, the SiCl₄ molecule has a tetrahedral shape.

**STUDY CHECK 10.5**
Use VSEPR theory to predict the shape of SCl₂.

**ANSWER**
The central atom S has four electron groups: two bonded atoms and two lone pairs of electrons. The shape of SCl₂ is bent, 109°.

**SAMPLE PROBLEM 10.6 Predicting Shape of an Ion**
Use VSEPR theory to predict the shape of the polyatomic ion NO₃⁻.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>NO₃⁻</td>
<td>shape</td>
<td>Lewis structure, electron groups, bonded atoms</td>
</tr>
</tbody>
</table>

**STEP 1** Draw the Lewis structure.

<table>
<thead>
<tr>
<th>Name of Element</th>
<th>Nitrogen</th>
<th>Oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>Symbol of Element</td>
<td>N</td>
<td>O</td>
</tr>
<tr>
<td>Atoms of Element</td>
<td>1</td>
<td>3</td>
</tr>
<tr>
<td>Valence Electrons</td>
<td>5 e⁻</td>
<td>6 e⁻</td>
</tr>
<tr>
<td>Total Electrons</td>
<td>1(5 e⁻) + 3(6 e⁻) + 1 e⁻ (charge) = 24 e⁻</td>
<td></td>
</tr>
</tbody>
</table>

The polyatomic ion NO₃⁻ contains three electron groups (two single bonds between the central N atom and O atoms, and one double bond between N and O). Note that the double bond can be drawn to any of the O atoms, which results in three resonance structures. However, we need just one of the structures to predict its shape.

\[
\text{NO}_3^- :\overset{\ddot{\text{N}}}{\text{O}}\equiv\overset{\ddot{\text{O}}}{\text{O}}\]

**STEP 2** Arrange the electron groups around the central atom to minimize repulsion. To minimize repulsion, the three electron groups have a trigonal planar geometry.

**STEP 3** Use the atoms bonded to the central atom to determine the shape.
Because NO₃⁻ has three bonded atoms and no lone pairs, it has a trigonal planar shape.

**STUDY CHECK 10.6**
Use VSEPR theory to predict the shape of ClO₂⁻ (Cl is the central atom).

**ANSWER**
The central atom Cl has four electron groups: two bonded atoms and two lone pairs of electrons. The shape of ClO₂⁻ is bent, 109°.
10.4 Electronegativity and Bond Polarity

LEARNING GOAL Use electronegativity to determine the polarity of a bond.

We can learn more about the chemistry of compounds by looking at how bonding electrons are shared between atoms. The bonding electrons are shared equally in a bond between identical nonmetal atoms. However, when a bond is between atoms of different elements, the electron pairs are usually shared unequally. Then the shared pairs of electrons are attracted to one atom in the bond more than the other.

The electronegativity of an atom is its ability to attract the shared electrons in a chemical bond. Nonmetals have higher electronegativities than do metals, because nonmetals have a greater attraction for electrons than metals. On the electronegativity scale, fluorine was assigned a value of 4.0, and the electronegativities for all other elements were determined relative to the attraction of fluorine for shared electrons. The nonmetal fluorine, which has the highest electronegativity (4.0), is located in the upper right corner of the periodic table. The metal cesium, which has the lowest electronegativity (0.7), is located in the lower left corner of the periodic table. The electronegativities for the representative elements are shown in Figure 10.2. Note that there are no electronegativity values for the noble gases because they do not typically form bonds. The electronegativity values for transition elements are also low, but we have not included them in our discussion.
The difference in the electronegativity values of two atoms can be used to predict the type of chemical bond, ionic or covalent, that forms. For the \( \text{H} - \text{H} \) bond, the electronegativity difference is zero \((2.1 - 2.1 = 0)\), which means the bonding electrons are shared equally. We illustrate this by drawing a symmetrical electron cloud around the H atoms. A bond between atoms with identical or very similar electronegativity values is a nonpolar covalent bond. However, when covalent bonds are between atoms with different electronegativity values, the electrons are shared unequally; the bond is a polar covalent bond. The electron cloud for a polar covalent bond is unsymmetrical. For the \( \text{H} - \text{Cl} \) bond, there is an electronegativity difference of \( 3.0(\text{Cl}) - 2.1(\text{H}) = 0.9 \), which means that the \( \text{H} - \text{Cl} \) bond is polar covalent (see FIGURE 10.3). When finding the electronegativity difference, the smaller electronegativity is always subtracted from the larger; thus, the difference is always a positive number.

**FIGURE 10.2** The electronegativity values of representative elements in Group 1A (1) to Group 7A (17), which indicate the ability of atoms to attract shared electrons, increase going across a period from left to right and decrease going down a group.

**FIGURE 10.3** In the nonpolar covalent bond of \( \text{H}_2 \), electrons are shared equally. In the polar covalent bond of \( \text{HCl} \), electrons are shared unequally.

**Polarity of Bonds**

The difference in the electronegativity values of two atoms can be used to predict the type of chemical bond, ionic or covalent, that forms. For the \( \text{H} - \text{H} \) bond, the electronegativity difference is zero \((2.1 - 2.1 = 0)\), which means the bonding electrons are shared equally. We illustrate this by drawing a symmetrical electron cloud around the H atoms. A bond between atoms with identical or very similar electronegativity values is a nonpolar covalent bond. However, when covalent bonds are between atoms with different electronegativity values, the electrons are shared unequally; the bond is a polar covalent bond. The electron cloud for a polar covalent bond is unsymmetrical. For the \( \text{H} - \text{Cl} \) bond, there is an electronegativity difference of \( 3.0(\text{Cl}) - 2.1(\text{H}) = 0.9 \), which means that the \( \text{H} - \text{Cl} \) bond is polar covalent (see FIGURE 10.3). When finding the electronegativity difference, the smaller electronegativity is always subtracted from the larger; thus, the difference is always a positive number.

**Dipoles and Bond Polarity**

The polarity of a bond depends on the difference in the electronegativity values of its atoms. In a polar covalent bond, the shared electrons are attracted to the more electronegative atom, which makes it partially negative, because of the negatively charged...
electrons around that atom. At the other end of the bond, the atom with the lower electronegativity becomes partially positive because of the lack of electrons at that atom.

A bond becomes more polar as the electronegativity difference increases. A polar covalent bond that has a separation of charges is called a dipole. The positive and negative ends of the dipole are indicated by the lowercase Greek letter delta with a positive or negative sign, \( \delta^+ \) and \( \delta^- \). Sometimes we use an arrow that points from the positive charge to the negative charge to indicate the dipole.

**Variations in Bonding**

The variations in bonding are continuous; there is no definite point at which one type of bond stops and the next starts. When the electronegativity difference is between 0.0 and 0.4, the electrons are considered to be shared equally in a nonpolar covalent bond. For example, the C—C bond (2.5 − 2.5 = 0.0) and the C—H bond (2.5 − 2.1 = 0.4) are classified as nonpolar covalent bonds.

As the electronegativity difference increases, the shared electrons are attracted more strongly to the more electronegative atom, which increases the polarity of the bond. When the electronegativity difference is from 0.5 to 1.8, the bond is a polar covalent bond. For example, the O—H bond (3.5 − 2.1 = 1.4) is classified as a polar covalent bond (see TABLE 10.6).

---

**TABLE 10.6** Electronegativity Differences and Types of Bonds

<table>
<thead>
<tr>
<th>Electronegativity Difference</th>
<th>Bond Type</th>
<th>Electron Bonding</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0 to 0.4</td>
<td>Nonpolar covalent</td>
<td>Electrons shared equally</td>
</tr>
<tr>
<td>0.5 to 1.8</td>
<td>Polar covalent</td>
<td>Electrons shared unequally</td>
</tr>
<tr>
<td>1.9 to 3.3</td>
<td>Ionic</td>
<td>Electrons transferred</td>
</tr>
</tbody>
</table>

When the electronegativity difference is greater than 1.8, electrons are transferred from one atom to another, which results in an ionic bond. For example, the electronegativity difference for the ionic compound NaCl is 3.0 − 0.9 = 2.1. Thus, for large differences in electronegativity, we would predict an ionic bond (see TABLE 10.7).

---

**TABLE 10.7** Predicting Bond Type from Electronegativity Differences

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Bond</th>
<th>Type of Electron Sharing</th>
<th>Electronegativity Difference</th>
<th>Type of Bond</th>
<th>Reason</th>
</tr>
</thead>
<tbody>
<tr>
<td>H(_2)</td>
<td>H—H</td>
<td>Shared equally</td>
<td>2.1 − 2.1 = 0.0</td>
<td>Nonpolar covalent</td>
<td>From 0.0 to 0.4</td>
</tr>
<tr>
<td>BrCl</td>
<td>Br—Cl</td>
<td>Shared about equally</td>
<td>3.0 − 2.8 = 0.2</td>
<td>Nonpolar covalent</td>
<td>From 0.0 to 0.4</td>
</tr>
<tr>
<td>HBr</td>
<td>H(^+)− Br(^−)</td>
<td>Shared unequally</td>
<td>2.8 − 2.1 = 0.7</td>
<td>Polar covalent</td>
<td>From 0.5 to 1.8</td>
</tr>
<tr>
<td>HCl</td>
<td>H(^+)− Cl(^−)</td>
<td>Shared unequally</td>
<td>3.0 − 2.1 = 0.9</td>
<td>Polar covalent</td>
<td>From 0.5 to 1.8</td>
</tr>
<tr>
<td>NaCl</td>
<td>Na(^+)Cl(^−)</td>
<td>Electron transfer</td>
<td>3.0 − 0.9 = 2.1</td>
<td>Ionic</td>
<td>From 1.9 to 3.3</td>
</tr>
<tr>
<td>MgO</td>
<td>Mg(^2+)O(^2−)</td>
<td>Electron transfer</td>
<td>3.5 − 1.2 = 2.3</td>
<td>Ionic</td>
<td>From 1.9 to 3.3</td>
</tr>
</tbody>
</table>

*Values are taken from Figure 10.2.*
SAMPLE PROBLEM 10.7 Bond Polarity

Using electronegativity values, classify each of the following bonds as nonpolar covalent, polar covalent, or ionic:

\[ \text{O} \rightarrow \text{K}, \quad \text{Cl} \rightarrow \text{As}, \quad \text{N} \rightarrow \text{N}, \quad \text{P} \rightarrow \text{Br} \]

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>bonds</td>
<td>type of bonds</td>
<td>electronegativity values</td>
</tr>
</tbody>
</table>

For each bond, we obtain the electronegativity values and calculate the difference in electronegativity.

<table>
<thead>
<tr>
<th>Bond</th>
<th>Electronegativity Difference</th>
<th>Type of Bond</th>
</tr>
</thead>
<tbody>
<tr>
<td>[ \text{O} \rightarrow \text{K} ]</td>
<td>[ 3.5 - 0.8 = 2.7 ]</td>
<td>Ionic</td>
</tr>
<tr>
<td>[ \text{Cl} \rightarrow \text{As} ]</td>
<td>[ 3.0 - 2.0 = 1.0 ]</td>
<td>Polar covalent</td>
</tr>
<tr>
<td>[ \text{N} \rightarrow \text{N} ]</td>
<td>[ 3.0 - 3.0 = 0.0 ]</td>
<td>Nonpolar covalent</td>
</tr>
<tr>
<td>[ \text{P} \rightarrow \text{Br} ]</td>
<td>[ 2.8 - 2.1 = 0.7 ]</td>
<td>Polar covalent</td>
</tr>
</tbody>
</table>

**STUDY CHECK 10.7**

Using electronegativity values, classify each of the following bonds as nonpolar covalent, polar covalent, or ionic:

a. \[ \text{P} \rightarrow \text{Cl} \]  
   b. \[ \text{Br} \rightarrow \text{Br} \]  
   c. \[ \text{Na} \rightarrow \text{O} \]

**ANSWER**  

a. polar covalent (0.9)  
   b. nonpolar covalent (0.0)  
   c. ionic (2.6)
10.5 Polarity of Molecules

LEARNING GOAL Use the three-dimensional structure of a molecule to classify it as polar or nonpolar.

We have seen that covalent bonds in molecules can be polar or nonpolar. Now we will look at how the bonds in a molecule and its shape determine whether that molecule is classified as polar or nonpolar.

Nonpolar Molecules

In a nonpolar molecule, all the bonds are nonpolar or the polar bonds cancel each other out. Molecules such as H₂, Cl₂, and CH₄ are nonpolar because they contain only nonpolar covalent bonds.

\[ \text{H–H} \quad \text{Cl–Cl} \quad \text{H–C–H} \]

Nonpolar molecules

A nonpolar molecule also occurs when polar bonds (dipoles) cancel each other because they are in a symmetrical arrangement. For example, CO₂, a linear molecule, contains two equal polar covalent bonds whose dipoles point in opposite directions. As a result, the dipoles cancel out, making a CO₂ molecule nonpolar.

Another example of a nonpolar molecule is the CCl₄ molecule, which has four polar bonds symmetrically arranged around the central C atom. Each of the C–Cl bonds has the same polarity, but because they have a tetrahedral arrangement, their opposing dipoles cancel out. As a result, a molecule of CCl₄ is nonpolar.

Polar Molecules

In a polar molecule, one end of the molecule is more negatively charged than the other end. Polarity in a molecule occurs when the dipoles from the individual polar bonds do not cancel each other. For example, HCl is a polar molecule because it has one covalent bond that is polar.

In molecules with two or more electron groups, the shape, such as bent or trigonal pyramidal, determines whether the dipoles cancel. For example, we have seen that H₂O has a bent shape. Thus, a water molecule is polar because the individual dipoles do not cancel.

\[ \text{H–O–H} \]

H₂O is a polar molecule because its dipoles do not cancel.

The NH₃ molecule has a tetrahedral electron-group geometry with three bonded atoms, which gives it a trigonal pyramidal shape. Thus, the NH₃ molecule is polar because the individual N–H dipoles do not cancel.

\[ \text{H–N–H} \]

NH₃ is a polar molecule because its dipoles do not cancel.

In the molecule CH₃F, the C–F bond is polar covalent, but the three C–H bonds are nonpolar covalent. Because there is only one dipole in CH₃F, CH₃F is a polar molecule.

\[ \text{H–C–F} \]

CH₃F is a polar molecule.
SAMPLE PROBLEM 10.8 Polarity of Molecules

Determine whether a molecule of OF₂ is polar or nonpolar.

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>OF₂</td>
<td>polarity</td>
<td>Lewis structure, bond polarity</td>
</tr>
</tbody>
</table>

STEP 1 Determine if the bonds are polar covalent or nonpolar covalent. From Figure 10.2, F and O have an electronegativity difference of 0.5 (4.0 – 3.5 = 0.5), which makes each of the O—F bonds polar covalent.

STEP 2 If the bonds are polar covalent, draw the Lewis structure and determine if the dipoles cancel. The Lewis structure for OF₂ has four electron groups and two bonded atoms. The molecule has a bent shape in which the dipoles of the O—F bonds do not cancel. The OF₂ molecule would be polar.

OF₂ is a polar molecule.

STUDY CHECK 10.8

Would a PCl₃ molecule be polar or nonpolar?

ANSWER

polar

QUESTIONS AND PROBLEMS

10.5 Polarity of Molecules

LEARNING GOAL Use the three-dimensional structure of a molecule to classify it as polar or nonpolar.

10.33 Why is F₂ a nonpolar molecule, but HF is a polar molecule?
10.34 Why is CCl₄ a nonpolar molecule, but PCl₃ is a polar molecule?
10.35 Identify each of the following molecules as polar or nonpolar:
   a. CS₂   b. NF₃   c. CHF₃   d. SO₃

10.36 Identify each of the following molecules as polar or nonpolar:
   a. SeF₂   b. PBr₃   c. SiF₄   d. SO₂

10.37 The molecule CO₂ is nonpolar, but CO is a polar molecule. Explain.

10.38 The molecules CH₄ and CH₃Cl both have tetrahedral shapes. Why is CH₄ nonpolar whereas CH₃Cl is polar?

10.6 Intermolecular Forces between Atoms or Molecules

LEARNING GOAL Describe the intermolecular forces between ions, polar covalent molecules, and nonpolar covalent molecules.

In gases, the interactions between particles are minimal, which allows gas molecules to move far apart from each other. In solids and liquids, there are sufficient interactions between the particles to hold them close together. Such differences in properties are
explained by looking at the various kinds of intermolecular forces between particles including dipole–dipole attractions, hydrogen bonding, dispersion forces, as well as ionic bonds.

**Ionic Bonds**

Ionic bonds are the strongest of the attractive forces found in compounds. Thus, most ionic compounds are solids at room temperature. The ionic compound sodium chloride, NaCl, melts at 801 °C. Large amounts of energy are needed to overcome the strong attractive forces between positive and negative ions and change solid sodium chloride to a liquid.

**Dipole–Dipole Attractions**

All polar molecules are attracted to each other by dipole–dipole attractions. Because polar molecules have dipoles, the positively charged end of the dipole in one molecule is attracted to the negatively charged end of the dipole in another molecule.

**Hydrogen Bonds**

Polar molecules containing hydrogen atoms bonded to highly electronegative atoms of nitrogen, oxygen, or fluorine form especially strong dipole–dipole attractions. This type of attraction, called a hydrogen bond, occurs between the partially positive hydrogen atom in one molecule and the partially negative nitrogen, oxygen, or fluorine atom in another molecule. Hydrogen bonds are the strongest type of attractive forces between polar covalent molecules. They are a major factor in the formation and structure of biological molecules such as proteins and DNA.

**Dispersion Forces**

Very weak attractions called dispersion forces are the only intermolecular forces that occur between nonpolar molecules. Usually, the electrons in a nonpolar covalent molecule are distributed symmetrically. However, the movement of the electrons may place more of them in one part of the molecule than another, which forms a temporary dipole. These temporary dipoles align the molecules so that the positive end of one molecule is attracted to the negative end of another molecule. Although dispersion forces are very weak, they make it possible for nonpolar molecules to form liquids and solids.

Nonpolar covalent molecules have weak attractions when they form temporary dipoles.

**Attractive Forces and Melting Points**

The melting point of a substance is related to the strength of the attractive forces between its particles. A compound with weak attractive forces, such as dispersion forces, has a low melting point because only a small amount of energy is needed to separate the molecules and form a liquid. A compound with dipole–dipole attractions requires more energy to break the attractive forces between the molecules. A compound that can form hydrogen bonds requires even more energy to overcome the attractive forces that exist between its molecules. Larger amounts of energy are needed to overcome the strong attractive forces between positive and negative ions and to melt an ionic solid. For example, the ionic solid MgF₂ melts at 1248 °C. Table 10.8 compares the melting points of some substances with various kinds of attractive forces. The various types of attractions between particles in solids and liquids are summarized in Table 10.9.

| Table 10.8 Melting Points of Selected Substances |
|---------------------------------|------------------|
| **Substance** | **Melting Point (°C)** |
| Ionic Bonds |  |
| MgF₂ | 1248 |
| NaCl | 801 |
| Hydrogen Bonds |  |
| H₂O | 0 |
| NH₃ | −78 |
| Dipole–Dipole Attractions |  |
| HI | −51 |
| HBr | −89 |
| HCl | −115 |
| Dispersion Forces |  |
| Br₂ | −7 |
| Cl₂ | −101 |
| F₂ | −220 |
**TABLE 10.9 Comparison of Bonding and Attractive Forces**

<table>
<thead>
<tr>
<th>Type of Force</th>
<th>Particle Arrangement</th>
<th>Example</th>
<th>Strength</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Between Atoms or Ions</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ionic bond</td>
<td><img src="image" alt="Ionic bond diagram" /></td>
<td>Na⁺Cl⁻</td>
<td>Strong</td>
</tr>
<tr>
<td>Covalent bond (X = nonmetal)</td>
<td><img src="image" alt="Covalent bond diagram" /></td>
<td>Cl—Cl</td>
<td></td>
</tr>
<tr>
<td><strong>Between Molecules</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrogen bond (X = N, O, or F)</td>
<td><img src="image" alt="Hydrogen bond diagram" /></td>
<td>H⁺⋯H₋</td>
<td></td>
</tr>
<tr>
<td>Dipole–dipole attractions (X and Y = nonmetals)</td>
<td><img src="image" alt="Dipole–dipole attraction diagram" /></td>
<td>H⁺⋯Cl⁻</td>
<td></td>
</tr>
<tr>
<td>Dispersion forces (temporary shift of electrons in nonpolar bonds)</td>
<td><img src="image" alt="Dispersion forces diagram" /></td>
<td>F⁻⋯F⁻</td>
<td>Weak</td>
</tr>
</tbody>
</table>

**ENGAGE**

Why does GeH₄ have a higher boiling point than CH₄?

**Size, Mass, and Melting and Boiling Points**

As the size and mass of similar types of molecular compounds increase, there are more electrons that produce stronger temporary dipoles. As the molar mass of similar compounds increases, the dispersion forces also increase due to the increase in the number of electrons. In general, larger nonpolar molecules with increased molar masses also have higher melting and boiling points.

**SAMPLE PROBLEM 10.9 Intermolecular Forces between Particles**

Indicate the major type of intermolecular forces—dipole–dipole attractions, hydrogen bonding, or dispersion forces—expected of each of the following:

a. HF  
b. Br₂  
c. PCl₃

**TRY IT FIRST**

**SOLUTION**

a. HF is a polar molecule that interacts with other HF molecules by hydrogen bonding.

b. Br₂ is nonpolar; only dispersion forces provide temporary intermolecular forces.

c. The polarity of PCl₃ molecules provides dipole–dipole attractions.

**STUDY CHECK 10.9**

Why is the melting point of H₂S lower than that of H₂O?

**ANSWER**

The intermolecular forces between H₂S molecules are dipole–dipole attractions, whereas the intermolecular forces between H₂O molecules are hydrogen bonds, which are stronger and require more energy to break.
**QUESTIONS AND PROBLEMS**

**10.6 Intermolecular Forces between Atoms or Molecules**

**LEARNING GOAL** Describe the intermolecular forces between ions, polar covalent molecules, and nonpolar covalent molecules.

10.39 Identify the major type of intermolecular forces between the particles of each of the following:
   a. BrF     b. KCl     c. NF₃     d. Cl₂

10.40 Identify the major type of intermolecular forces between the particles of each of the following:
   a. HCl     b. MgF₂     c. PBr₃     d. NH₃

10.41 Identify the strongest intermolecular forces between the particles of each of the following:
   a. CH₃OH     b. CO     c. CF₄     d. CH₃—CH₃

10.42 Identify the strongest intermolecular forces between the particles of each of the following:
   a. O₂     b. SiH₄     c. CH₃Cl     d. H₂O₂

10.43 Identify the substance in each of the following pairs that would have the higher boiling point and explain your choice:
   a. HF or HBr     b. HF or NaF
   c. MgBr₂ or PBr₃     d. CH₄ or CH₃OH

10.44 Identify the substance in each of the following pairs that would have the higher boiling point and explain your choice:
   a. NaCl or HCl     b. H₂O or H₂Se
   c. NH₃ or PH₃     d. F₂ or HF

**10.7 Changes of State**

**LEARNING GOAL** Describe the changes of state between solids, liquids, and gases; calculate the energy involved.

The states and properties of gases, liquids, and solids depend on the types of attractive forces between their particles. Matter undergoes a **change of state** when it is converted from one state to another state (see **FIGURE 10.4**).

![FIGURE 10.4 ▶ Changes of state include melting and freezing, vaporization and condensation, sublimation and deposition.
Is heat added or released when liquid water freezes?](image-url)
Melting and Freezing

When heat is added to a solid, the particles move faster. At a temperature called the melting point (mp), the particles of a solid gain sufficient energy to overcome the attractive forces that hold them together. The particles in the solid separate and move about in random patterns. The substance is melting, changing from a solid to a liquid.

If the temperature of a liquid is lowered, the reverse process takes place. Kinetic energy is lost, the particles slow down, and attractive forces pull the particles close together. The substance is freezing. A liquid changes to a solid at the freezing point (fp), which is the same temperature as its melting point. Every substance has its own freezing (melting) point: Solid water (ice) melts at 0 °C when heat is added, and freezes at 0 °C when heat is removed. Gold melts at 1064 °C when heat is added, and freezes at 1064 °C when heat is removed.

During a change of state, the temperature of a substance remains constant. Suppose we have a glass containing ice and water. The ice melts when heat is added at 0 °C, forming liquid. When heat is removed at 0 °C, the liquid water freezes, forming solid. The process of melting requires heat; the process of freezing releases heat. Melting and freezing are reversible at 0 °C.

Heat of Fusion

During melting, the heat of fusion is the energy that must be added to convert exactly 1 g of solid to liquid at the melting point. For example, 334 J of heat is needed to melt exactly 1 g of ice at its melting point (0 °C).

\[ \text{H}_2\text{O}(s) + 334 \text{ J/g} \rightarrow \text{H}_2\text{O}(l) \quad \text{Endothermic} \]

Heat of Fusion for Water

\[ \frac{334 \text{ J}}{1 \text{ g H}_2\text{O}} \quad \text{and} \quad \frac{1 \text{ g H}_2\text{O}}{334 \text{ J}} \]

The heat of fusion is also the quantity of heat that must be removed to freeze exactly 1 g of water at its freezing point (0 °C).

Water is sometimes sprayed in fruit orchards during subfreezing weather. If the air temperature drops to 0 °C, the water begins to freeze. Because heat is released as the water molecules form solid ice, the air warms above 0 °C and protects the fruit.

\[ \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s) + 334 \text{ J/g} \quad \text{Exothermic} \]

The heat of fusion can be used as a conversion factor in calculations. For example, to determine the heat needed to melt a sample of ice, the mass of ice, in grams, is multiplied by the heat of fusion. Because the temperature remains constant as long as the ice is melting, there is no temperature change given in the calculation, as shown in Sample Problem 10.10.

Calculating Heat to Melt (or Freeze) Water

Heat = mass × heat of fusion

\[ J = g \times \frac{334 \text{ J}}{g} \]

SAMPLE PROBLEM 10.10 Heat of Fusion

Ice bags are used by sports trainers to treat muscle injuries. If 260 g of ice are placed in an ice bag, how much heat, in joules, will be absorbed to melt all the ice at 0 °C?

TRY IT FIRST
**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>260.0 g of ice at 0 °C</td>
<td>joules to melt ice at 0 °C</td>
<td>heat of fusion</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to convert the given quantity to the needed quantity.

grams of ice  \[ \text{Heat of fusion} \] joules

**STEP 3** Write the heat conversion factor.

\[
\frac{1 \text{ g of H}_2\text{O} (s \rightarrow l) = 334 \text{ J}}{1 \text{ g H}_2\text{O}} \quad \text{and} \quad \frac{1 \text{ g H}_2\text{O}}{334 \text{ J}}
\]

**STEP 4** Set up the problem and calculate the needed quantity.

Three SFs

\[
26.0 \text{ g } \text{H}_2\text{O} \times \frac{334 \text{ J}}{1 \text{ g H}_2\text{O}} = 8680 \text{ J (8.68 } \times 10^3 \text{ J)}
\]

Three SFs  Exact  Three SFs

**STUDY CHECK 10.10**

In a freezer, 125 g of water at 0 °C is placed in an ice cube tray. How much heat, in kilojoules, must be removed to form ice cubes at 0 °C?

**ANSWER**

41.8 kJ
Evaporation, Boiling, and Condensation

Water in a mud puddle disappears, unwrapped food dries out, and clothes hung on a line dry. **Evaporation** is taking place as water molecules with sufficient energy escape from the liquid surface and enter the gas phase (see **Figure 10.5a**). The loss of the “hot” water molecules removes heat, which cools the remaining liquid water. As heat is added, more and more water molecules gain sufficient energy to evaporate. At the **boiling point (bp)**, the molecules within a liquid have enough energy to overcome their attractive forces and become a gas. We observe the **boiling** of a liquid such as water as gas bubbles form throughout the liquid, rise to the surface, and escape (see **Figure 10.5b**).

When heat is removed, a reverse process takes place. In **condensation**, water vapor is converted to liquid as the water molecules lose kinetic energy and slow down. Condensation occurs at the same temperature as boiling but differs because heat is removed. You may have noticed that condensation occurs when you take a hot shower and the water vapor forms water droplets on a mirror. Because a substance loses heat as it condenses, its surroundings become warmer. That is why, when a rainstorm is approaching, you may notice a warming of the air as gaseous water molecules condense to rain.

**Heat of Vaporization and Condensation**

The **heat of vaporization** is the energy that must be added to convert exactly 1 g of liquid to gas at its boiling point. For water, 2260 J is needed to convert 1 g of water to vapor at 100 °C.

\[
\text{H}_2\text{O}(l) + 2260 \text{ J/g} \rightarrow \text{H}_2\text{O}(g) \quad \text{Endothermic}
\]

This same amount of heat is released when 1 g of water vapor (gas) changes to liquid at 100 °C.

\[
\text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(l) + 2260 \text{ J/g} \quad \text{Exothermic}
\]

Therefore, 2260 J/g is also the **heat of condensation** of water.

**Heat of Vaporization for Water**

\[
\frac{2260 \text{ J}}{1 \text{ g H}_2\text{O}} \quad \text{and} \quad \frac{1 \text{ g H}_2\text{O}}{2260 \text{ J}}
\]

To determine the heat needed to boil a sample of water, the mass, in grams, is multiplied by the heat of vaporization. Because the temperature remains constant as long as the water is boiling, there is no temperature change given in the calculation.

**Calculating Heat to Vaporize (or Condense) Water**

\[
\text{Heat} = \text{mass} \times \text{heat of vaporization}
\]

\[
J = g \times \frac{2260 \text{ J}}{g}
\]

Just as substances have different melting and boiling points, they also have different heats of fusion and vaporization. As seen in **Figure 10.6**, the heats of vaporization are larger than the heats of fusion.

**Heating and Cooling Curves**

All the changes of state during the heating or cooling of a substance can be illustrated visually. On a **heating curve** or **cooling curve**, the temperature is shown on the vertical axis and the loss or gain of heat is shown on the horizontal axis.
A 5.0-g sample of ammonia has a boiling point of \(-33 \, ^\circ C\) and a heat of vaporization of 1380 J/g. The heat change when the sample is completely vaporized is 6900 J. Explain why the temperature \(-33 \, ^\circ C\) was not used in the calculation of the heat.

**FIGURE 10.6** For any substance, the heat of vaporization is greater than the heat of fusion.

**Q** Why does the formation of a gas require more energy than the formation of a liquid of the same compound?

### Steps on a Heating Curve

The first diagonal line indicates a warming of a solid as heat is added. When the melting temperature is reached, a horizontal line, or plateau, indicates that the solid is melting. As melting takes place, the solid is changing to liquid without any change in temperature (see **FIGURE 10.7**).

Once all the particles are in the liquid state, heat that is added will increase the temperature of the liquid. This increase is drawn as a diagonal line from the melting point temperature to the boiling point temperature. Once the liquid reaches its boiling point, a horizontal line indicates that the temperature is constant as liquid changes to gas. Because the heat of vaporization is greater than the heat of fusion, the horizontal line at the boiling point is longer than the line at the melting point. Once all the liquid becomes gas, adding more heat increases the temperature of the gas.
Steps on a Cooling Curve

A cooling curve is a diagram of the cooling process in which the temperature decreases as heat is removed. Initially, a diagonal line to the boiling (condensation) point is drawn to show that heat is removed from a gas until it begins to condense. At the boiling (condensation) point, a horizontal line is drawn that indicates a change of state as gas condenses to form a liquid. After all the gas has changed into liquid, further cooling lowers the temperature. The decrease in temperature is shown as a diagonal line from the condensation temperature to the freezing temperature. At the freezing point, another horizontal line indicates that liquid is changing to solid at the freezing point temperature. Once all the substance is frozen, the removal of more heat decreases the temperature of the solid below its freezing point, which is shown as a diagonal line below its freezing point.

Combining Energy Calculations

Up to now, we have calculated one step in a heating or cooling curve. However, a problem may require a combination of steps in which a substance may change state and then the new state undergoes a temperature change followed by another change of state. When the temperature of a substance is changing and not its state, we need to use the heat equation, which includes the $\Delta T$.

$$\text{Heat (J)} = \text{mass (g)} \times \text{temperature change (} \Delta T \text{)} \times \text{specific heat (J/g °C)}$$

The heat is calculated for each step separately, and the results are added together to give the total energy, as seen in Sample Problem 10.11.

**SAMPLE PROBLEM 10.11 Combining Heat Calculations**

Calculate the total heat, in joules, needed to convert 15.0 g of liquid ethanol at 25.0 °C to gas at its boiling point, 78.0 °C. The specific heat of ethanol is 2.46 J/g °C. The heat of vaporization for liquid ethanol is 841 J/g.

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>15.0 g of ethanol, temperature change (25.0 °C to 78.0 °C), specific heat of ethanol = 2.46 J/g °C, heat of vaporization of ethanol = 841 J/g</td>
<td>total joules to raise temperature and convert to gas</td>
<td>combine heat from temperature change and change of state (heat of vaporization)</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to convert the given quantity to the needed quantity.

When several changes occur, draw a diagram of heating and changes of state.

Total joules =

$$(1) \text{ joules to heat ethanol from } 25.0 \degree \text{C} \rightarrow 78.0 \degree \text{C}$$

$$\text{Joules} = \text{mass (g)} \times \text{temperature change (} \Delta T \text{)} \times \text{specific heat (J/g °C)}$$
(2) joules to change liquid ethanol to gas at 78.0 °C

\[\text{Joules} = \text{mass (g)} \times \text{heat of vaporization (J/g)}\]

\[
\begin{align*}
T^\circ C & \uparrow \\
78.0 & \quad 2 \quad \text{Liquid } \rightarrow \text{Gas} \\
25.0 & \quad 1 \quad \text{Liquid}
\end{align*}
\]

Heat Added

**STEP 3** Write the heat conversion factors needed.

\[
SH_{\text{Ethanol}} = \frac{2.46 \text{ J}}{\text{g} \cdot ^\circ C} \quad \text{and} \quad \frac{2.46 \text{ J}}{2.46 \text{ J}} = \frac{1 \text{ g of ethanol (l \rightarrow g)}}{1 \text{ g ethanol}} = \frac{841 \text{ J}}{841 \text{ J}}
\]

**STEP 4** Set up the problem and calculate the needed quantity.

\[\Delta T = 78.0^\circ C - 25.0^\circ C = 53.0^\circ C\]

**Temperature change:** Heat needed to warm ethanol (liquid) from 25.0 °C to 78.0 °C.

\[
15.0 \text{ g} \times 53.0^\circ C \times \frac{2.46 \text{ J}}{1 \text{ g} \cdot ^\circ C} = 1960 \text{ J}
\]

**Change of state at constant temperature:** Heat needed to convert ethanol (liquid) to ethanol (gas) at 78.0 °C.

\[
15.0 \text{ g ethanol} \times \frac{841 \text{ J}}{1 \text{ g ethanol}} = 12600 \text{ J}
\]

**Calculate the total heat:**

- Heating ethanol (25.0 °C to 78.0 °C) 1960 J
- Changing liquid to gas (78.0 °C) 12600 J
- Total heat needed 14600 J (rounded off)

**STUDY CHECK 10.11**

How many kilojoules are released when 75.0 g of ethanol at 78 °C condenses, cools to −114 °C, and freezes? The specific heat of ethanol is 2.46 J/g °C. For heats of fusion and vaporization, see Figure 10.6. (Hint: The solution will require three energy calculations.)

**ANSWER**

106.7 kJ
Sublimation and Deposition

In a process called sublimation, the particles on the surface of a solid change directly to a gas with no temperature change and without going through the liquid state. In the reverse process called deposition, gas particles change directly to a solid. For example, dry ice, which is solid carbon dioxide, sublimes at −78 °C. It is called “dry” because it does not form a liquid as it warms. In extremely cold areas, snow does not melt but sublimes directly to water vapor.

When frozen foods are left in the freezer for a long time, so much water sublimes that foods, especially meats, become dry and shrunk, a condition called freezer burn. Deposition occurs in a freezer when water vapor forms ice crystals on the surface of freezer bags and frozen food.

Freeze-dried foods prepared by sublimation are convenient for long-term storage and for camping and hiking. A food that has been frozen is placed in a vacuum chamber where it dries as the ice sublimes. The dried food retains all of its nutritional value and needs only water to be edible. A food that is freeze-dried does not need refrigeration because bacteria cannot grow without moisture.

### CHEMISTRY LINK TO HEALTH

Steam Burns

Hot water at 100 °C will cause burns and damage to the skin. However, getting steam on the skin is even more dangerous. If 25 g of hot water at 100 °C falls on a person’s skin, the temperature of the water will drop to body temperature, 37 °C. The heat released during cooling can cause severe burns. The amount of heat can be calculated from the mass, the temperature change, 100 °C − 37 °C = 63 °C, and the specific heat of water, 4.184 J/g °C.

\[
25 \text{ g} \times 63 \text{ °C} \times \frac{4.184 \text{ J}}{\text{g} \text{ °C}} = 6600 \text{ J of heat released when water cools}
\]

However, getting steam on the skin is even more damaging. The condensation of the same quantity of steam to liquid at 100 °C releases much more heat—almost ten times as much. This amount of heat can be calculated using the heat of vaporization, which is 2260 J/g for water at 100 °C.

\[
25 \text{ g} \times \frac{2260 \text{ J}}{1 \text{ g}} = 57 \text{ 000 J released when water (gas) condenses to water (liquid) at 100 °C}
\]

The total heat released is calculated by combining the heat from the condensation at 100 °C and the heat from cooling of the steam from 100 °C to 37 °C (body temperature). We can see that most of the heat is from the condensation of steam. This large amount of heat released on the skin is what causes damage from steam burns.

\[
\text{Condensation (100 °C)} = 57 \text{ 000 J}
\]
\[
\text{Cooling (100 °C to 37 °C)} = 6600 \text{ J}
\]
\[
\text{Heat released} = 64 \text{ 000 J (rounded off)}
\]

The amount of heat released from steam is almost ten times greater than the heat from the same amount of hot water.
LEARNING GOAL Describe the changes of state between solids, liquids, and gases; calculate the energy involved.

10.45 Using Figure 10.6, calculate the heat change needed for each of the following at the melting/freezing point:
   a. joules to melt 65.0 g of ice
   b. joules to melt 17.0 g of benzene
   c. kilojoules to freeze 225 g of acetone
   d. kilojoules to freeze 0.0500 kg of water

10.46 Using Figure 10.6, calculate the heat change needed for each of the following at the boiling/condensation point:
   a. joules to vaporize 10.0 g of water
   b. kilojoules to vaporize 50.0 g of ethanol
   c. joules to condense 8.00 g of acetic acid
   d. kilojoules to condense 0.175 kg of ammonia

10.47 Using Figure 10.6, calculate the heat change needed for each of the following at the boiling/condensation point:
   a. joules to vaporize 10.0 g of water
   b. kilojoules to vaporize 275 g of water
   c. kilojoules to melt 145 g of ammonia
   d. kilojoules to melt 5.00 kg of water

10.48 Using Figure 10.6, calculate the heat change needed for each of the following at the boiling/condensation point:
   a. joules to condense 10.0 g of steam
   b. kilojoules to condense 76.0 g of acetic acid
   c. joules to vaporize 44.0 g of ammonia
   d. kilojoules to vaporize 5.0 kg of water

10.49 Using Figure 10.6 and the specific heat of water, 4.184 J/g °C, calculate the total amount of heat for each of the following:
   a. joules needed to warm 20.0 g of water at 15 °C to 72 °C
   b. joules needed to melt 50.0 g of ice at 0.0 °C and to warm the liquid to 55.0 °C
   c. kilojoules released when 15.0 g of steam condenses at 100 °C and the liquid cools to 0 °C
   d. kilojoules needed to melt 24.0 g of ice at 0 °C, warm the liquid to 100 °C, and change it to steam at 100 °C

10.50 Using Figure 10.6 and the specific heat of water, 4.184 J/g °C, calculate the total amount of heat for each of the following:
   a. joules to condense 125 g of steam at 100 °C and to cool the liquid to 15.0 °C
   b. joules needed to melt a 525-g ice sculpture at 0 °C and to warm the liquid to 15.0 °C
   c. kilojoules released when 85.0 g of steam condenses at 100 °C, the liquid cools to 0 °C, and freezes
   d. joules to warm 55.0 mL of water (density = 1.00 g/mL) from 10.0 °C to 100 °C and vaporize it at 100 °C

Applications

10.51 An ice bag containing 275 g of ice at 0 °C was used to treat sore muscles. When the bag was removed, the ice had melted and the liquid water had a temperature of 24.0 °C. How many kilojoules of heat were absorbed?

10.52 A patient arrives in the emergency room with a burn caused by steam. Calculate the heat, in kilojoules, that is released when 18.0 g of steam at 100 °C hits the skin, condenses, and cools to body temperature at 37.0 °C.

10.53 At the boiling point of nitrogen, N₂, how many joules are needed to vaporize 15.8 g of N₂?

10.54 In the preparation of liquid nitrogen, how many kilojoules must be removed from 112 g of N₂ gas at its boiling point to form liquid N₂?

10.55 Using the electronegativity values in Figure 10.2, for each of the following bonds found in eosin, calculate the electronegativity difference and predict if the bond is nonpolar covalent, polar covalent, or ionic:
   a. C—C
   b. C—H
   c. C—O

10.56 Using the electronegativity values in Figure 10.2, for each of the following bonds found in eosin, calculate the electronegativity difference and predict if the bond is nonpolar covalent, polar covalent, or ionic:
   a. O—H
   b. Na—O
   c. C—Br

10.57 a. Draw the resonance structures for carbonate (CO₃²⁻).
b. Determine the shape of the carbonate ion.

10.58 a. Draw the resonance structures for bicarbonate (HCO₃⁻) with a central C atom and an H atom attached to one of the O atoms.
b. Determine the shape of the bicarbonate ion around the C atom.
**CHAPTER REVIEW**

**10.1 Lewis Structures for Molecules and Polyatomic Ions**

**LEARNING GOAL** Draw the Lewis structures for molecular compounds or polyatomic ions with single and multiple bonds.

- The total number of valence electrons is determined for all the atoms in the molecule or polyatomic ion.
- Any negative charge is added to the total valence electrons, whereas any positive charge is subtracted.
- In the Lewis structure, a bonding pair of electrons is placed between the central atom and each of the attached atoms.
- Any remaining valence electrons are used as lone pairs to complete the octets of the surrounding atoms and then the central atom.
- If octets are not completed, one or more lone pairs of electrons are placed as bonding pairs forming double or triple bonds.

**10.2 Resonance Structures**

**LEARNING GOAL** Draw Lewis structures for molecules or polyatomic ions that have two or more resonance structures.

- Resonance structures can be drawn when there are two or more equivalent Lewis structures for a molecule or ion with multiple bonds.

**10.3 Shapes of Molecules and Polyatomic Ions (VSEPR Theory)**

**LEARNING GOAL** Predict the three-dimensional structure of a molecule or a polyatomic ion.

- The shape of a molecule is determined from the Lewis structure, the electron-group geometry, and the number of bonded atoms.
- The electron-group geometry around a central atom with two electron groups is linear; with three electron groups, the geometry is trigonal planar; and with four electron groups, the geometry is tetrahedral.
- When all the electron groups are bonded to atoms, the shape has the same name as the electron arrangement.
- A central atom with three electron groups and two bonded atoms has a bent shape, 120°.
- A central atom with four electron groups and three bonded atoms has a trigonal pyramidal shape.
- A central atom with four electron groups and two bonded atoms has a bent shape, 109°.

**10.4 Electronegativity and Bond Polarity**

**LEARNING GOAL** Use electronegativity to determine the polarity of a bond.

- Electronegativity is the ability of an atom to attract the electrons it shares with another atom. In general, the electronegativities of metals are low, whereas nonmetals have high electronegativities.
In a nonpolar covalent bond, atoms share electrons equally.

In a polar covalent bond, the electrons are unequally shared because they are attracted to the more electronegative atom.

The atom in a polar bond with the lower electronegativity is partially positive (δ+), and the atom with the higher electronegativity is partially negative (δ−).

Atoms that form ionic bonds have large differences in electronegativities.

10.5 Polarity of Molecules

LEARNING GOAL Use the three-dimensional structure of a molecule to classify it as polar or nonpolar.

- Nonpolar molecules contain nonpolar covalent bonds or have an arrangement of bonded atoms that causes the dipoles to cancel out.
- In polar molecules, the dipoles do not cancel.

10.6 Intermolecular Forces Between Atoms or Molecules

LEARNING GOAL Describe the intermolecular forces between ions, polar covalent molecules, and nonpolar covalent molecules.

- In ionic solids, oppositely charged ions are held in a rigid structure by ionic bonds.
- Intermolecular forces called dipole–dipole attractions and hydrogen bonds hold the solid and liquid states of polar molecular compounds together.
- Nonpolar compounds form solids and liquids by weak attractions between temporary dipoles called dispersion forces.

10.7 Changes of State

LEARNING GOAL Describe the changes of state between solids, liquids, and gases; calculate the energy involved.

- Melting occurs when the particles in a solid absorb enough energy to break apart and form a liquid.
- The amount of energy required to convert exactly 1 g of solid to liquid is called the heat of fusion.
- For water, 334 J is needed to melt 1 g of ice or must be removed to freeze 1 g of water.
- Evaporation occurs when particles in a liquid state absorb enough energy to break apart and form gaseous particles.
- Boiling is the vaporization of liquid at its boiling point. The heat of vaporization is the amount of heat needed to convert exactly 1 g of liquid to vapor.
- For water, 2260 J is needed to vaporize 1 g of water or must be removed to condense 1 g of steam.
- A heating or cooling curve illustrates the changes in temperature and state as heat is added to or removed from a substance. Plateaus on the graph indicate changes of state.
- The total heat absorbed or removed from a substance undergoing temperature changes and changes of state is the sum of energy calculations for changes of state and changes in temperature.
- Sublimation is a process whereby a solid changes directly to a gas.

**KEY TERMS**

- **bent** The shape of a molecule with two bonded atoms and one lone pair or two lone pairs.
- **boiling** The formation of bubbles of gas throughout a liquid.
- **boiling point (bp)** The temperature at which a liquid changes to gas (boils) and gas changes to liquid (condenses).
- **change of state** The transformation of one state of matter to another, for example, solid to liquid, liquid to solid, liquid to gas.
- **condensation** The change of state from a gas to a liquid.
- **cooling curve** A diagram that illustrates temperature changes and changes of state for a substance as heat is removed.
- **deposition** The change of a gas directly to a solid; the reverse of sublimation.
- **dipole** The separation of positive and negative charge in a polar bond indicated by an arrow that is drawn from the more positive atom to the more negative atom.
- **dipole–dipole attractions** Attractive forces between oppositely charged ends of polar molecules.
- **dispersion forces** Weak dipole bonding that results from a momentary polarization of nonpolar molecules.
- **double bond** A sharing of two pairs of electrons by two atoms.
- **electronegativity** The relative ability of an element to attract electrons in a bond.
- **evaporation** The formation of a gas (vapor) by the escape of high-energy molecules from the surface of a liquid.
- **freezing** A change of state from liquid to solid.
- **freezing point (fp)** The temperature at which a liquid changes to a solid (freezes) and a solid changes to a liquid (melts).
- **heat of fusion** The energy required to melt exactly 1 g of a substance at its melting point. For water, 334 J is needed to melt 1 g of ice; 334 J is released when 1 g of water freezes.
- **heat of vaporization** The energy required to vaporize 1 g of a substance at its boiling point. For water, 2260 J is needed to vaporize exactly 1 g of water; 1 g of steam gives off 2260 J when it condenses.
- **heating curve** A diagram that illustrates the temperature changes and changes of state of a substance as it is heated.
- **hydrogen bond** The attraction between a partially positive H atom and a strongly electronegative atom of N, O, or F.
- **Lewis structure** A structure drawn in which the valence electrons of all the atoms are arranged to give octets except two electrons for hydrogen.
- **linear** The shape of a molecule that has two bonded atoms and no lone pairs.
- **melting** The change of state from a solid to a liquid.
- **melting point (mp)** The temperature at which a solid becomes a liquid (melts). It is the same temperature as the freezing point.
- **nonpolar covalent bond** A covalent bond in which the electrons are shared equally between atoms.
- **nonpolar molecule** A molecule that has only nonpolar bonds or in which the bond dipoles cancel.
polar covalent bond A covalent bond in which the electrons are shared unequally between atoms.

polar molecule A molecule containing bond dipoles that do not cancel.

polarity A measure of the unequal sharing of electrons, indicated by the difference in electronegativities.

resonance structures Two or more Lewis structures that can be drawn for a molecule or polyatomic ion by placing a multiple bond between different atoms.

sublimation The change of state in which a solid is transformed directly to a gas without forming a liquid first.

CORE CHEMISTRY SKILLS

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

Drawing Lewis Structures (10.1)

- The Lewis structure for a molecule or polyatomic ion shows the sequence of atoms, the bonding pairs of electrons shared between atoms, and the nonbonding or lone pairs of electrons.
- Double or triple bonds result when a second or third electron pair is shared between the same atoms to complete octets.

Example: Draw the Lewis structure for CS₂.

Answer: The central atom in the atom arrangement is C.

\[
\text{S} \quad \text{C} \quad \text{S}
\]

Determine the total number of valence electrons.

\[
2 \text{S} \times 6e^- = 12e^- \\
1 \text{C} \times 4e^- = 4e^- \\
\text{Total} = 16e^-
\]

Attach each bonded atom to the central atom using a pair of electrons. Two bonding pairs use four electrons.

\[
\text{S} - \text{C} - \text{S}
\]

Place the 12 remaining electrons as lone pairs around the S atoms.

\[
\vdash \text{S} - \text{C} - \text{S} \vdash
\]

To complete the octet for C, a lone pair of electrons from each of the S atoms is shared with C, which forms two double bonds.

\[
\vdash \text{S} - \text{C} - \text{S} \vdash \text{ or } \vdash \text{S} = \text{C} = \text{S} \vdash
\]

Drawing Resonance Structures (10.2)

- When a molecule or polyatomic ion contains multiple bonds, it may be possible to draw more than one Lewis structure called resonance structures.

Example: Draw the resonance structures for NO₂⁻.

Answer: The central atom in the atom arrangement is N.

\[
\text{O} \quad \text{N} \quad \text{O}
\]

Determine the total number of valence electrons.

\[
1 \text{N} \times 5e^- = 5e^- \\
2 \text{O} \times 6e^- = 12e^- \\
\text{negative charge} = 1e^- \\
\text{Total} = 18e^-
\]

Use electron pairs to attach each bonded atom to the central atom. Two bonding pairs use four electrons.

\[
\text{O} - \text{N} - \text{O}
\]

Place the 14 remaining electrons as lone pairs around the O and N atoms.

\[
\begin{align*}
\vdash \text{O} \vdash & \text{ or } \vdash \text{O} \vdash \\
\end{align*}
\]

To complete the octet for N, a lone pair from one O atom is shared with N, which forms one double bond. Because there are two O atoms, there are two resonance structures.

\[
\begin{align*}
\vdash \text{O} = \text{N} - \text{O} \vdash & \text{ or } \vdash \text{O} = \text{N} - \text{O} \vdash \\
\end{align*}
\]

Predicting Shape (10.3)

- The three-dimensional shape of a molecule or polyatomic ion is determined by drawing a Lewis structure and identifying the number of electron groups (one or more electron pairs) around the central atom and the number of bonded atoms.
- In the valence shell electron-pair repulsion (VSEPR) theory, the electron groups are arranged as far apart as possible around a central atom to minimize the repulsion.
- A central atom with two electron groups bonded to two atoms is linear. A central atom with three electron groups bonded to three atoms is trigonal planar, and to two atoms is bent (120°). A central atom with four electron groups bonded to four atoms is tetrahedral, to three atoms is trigonal pyramidal, and to two atoms is bent (109°).

Example: Predict the shape for NO₂⁻.

Answer: Using the resonance structures we drew for NO₂⁻, we count three electron groups around the central N atom: one double bond, one single bond, and a lone pair of electrons.

\[
\begin{align*}
\vdash \text{O} = \text{N} - \text{O} \vdash & \text{ or } \vdash \text{O} = \text{N} - \text{O} \vdash \\
\end{align*}
\]

The electron-group geometry is trigonal planar, but with the central N atom bonded to two O atoms, the shape is bent, 120°.
Using Electronegativity (10.4)

- The electronegativity values indicate the ability of atoms to attract shared electrons.
- Electronegativity values increase going across a period from left to right, and decrease going down a group.
- A nonpolar covalent bond occurs between atoms with identical or very similar electronegativity values such that the electronegativity difference is 0.0 to 0.4.
- A polar covalent bond typically occurs when electrons are shared unequally between atoms with electronegativity differences from 0.5 to 1.8.
- An ionic bond typically occurs when the difference in electronegativity is greater than 1.8.

Example: Use electronegativity values to classify each of the following bonds as nonpolar covalent, polar covalent, or ionic:

- a. Sr—Cl
- b. C—S
- c. O—Br

Answer: a. An electronegativity difference of 2.0 (Cl 3.0 — Sr 1.0) makes this an ionic bond.
- b. An electronegativity difference of 0.0 (C 2.5 — S 2.5) makes this a nonpolar covalent bond.
- c. An electronegativity difference of 0.7 (O 3.5 — Br 2.8) makes this a polar covalent bond.

Identifying Polarity of Molecules (10.5)

- A molecule is nonpolar if all of its bonds are nonpolar or it has polar bonds that cancel out. CCl₄ is a nonpolar molecule that consists of four polar bonds that cancel out.

- A molecule is polar if it contains polar bonds that do not cancel out. H₂O is a polar molecule that consists of polar bonds that do not cancel out.

Example: Predict whether AsCl₃ is polar or nonpolar.

Answer: From its Lewis structure, we see that AsCl₃ has four electron groups with three bonded atoms.

The shape of a molecule of AsCl₃ would be trigonal pyramidal with three polar bonds (As—Cl = 3.0 — 2.0 = 1.0) that do not cancel. Thus, it is a polar molecule.

Identifying Intermolecular Forces (10.6)

- Dipole–dipole attractions occur between the dipoles in polar compounds because the positively charged end of one molecule is attracted to the negatively charged end of another molecule.

- Strong dipole–dipole attractions called hydrogen bonds occur in compounds in which H is bonded to N, O, or F. The partially positive H atom in one molecule has a strong attraction to the partially negative N, O, or F in another molecule.

- Dispersion forces are very weak attractive forces between nonpolar molecules that occur when temporary dipoles form as electrons are unsymmetrically distributed.

Example: Identify the strongest type of intermolecular forces in each of the following:

- a. HF
- b. F₂
- c. NF₃

Answer: a. HF molecules, which are polar with H bonded to F, have hydrogen bonding.
- b. Nonpolar F₂ molecules have only dispersion forces.
- c. NF₃ molecules, which are polar, have dipole–dipole attractions.

- Substances with ionic bonds have the highest melting and boiling points. Substances with hydrogen bonds have higher melting and boiling points than compounds with only dipole–dipole attractions. Substances with only dispersion forces would typically have the lowest melting and boiling points, which increase as molar mass increases.

Example: Identify the compound with the highest boiling point in each of the following:

- a. HI, HBr, HF
- b. F₂, Cl₂, I₂

Answer: a. HBr and HI have dipole–dipole attractions, but HF has hydrogen bonding, which gives HF the highest boiling point.
- b. Because F₂, Cl₂, and I₂ have only dispersion forces, I₂ with the greatest molar mass would have the highest boiling point.

Calculating Heat for Change of State (10.7)

- At the melting/freezing point, the heat of fusion is absorbed/released to convert 1 g of a solid to a liquid or 1 g of liquid to a solid.

- For example, 334 J of heat is needed to melt (freeze) exactly 1 g of ice at its melting (freezing) point (0 °C).

- At the boiling/condensation point, the heat of vaporization is absorbed/released to convert exactly 1 g of liquid to gas or 1 g of gas to liquid.

- For example, 2260 J of heat is needed to boil (condense) exactly 1 g of water/steam at its boiling (condensation) point, 100 °C.

Example: What is the quantity of heat, in kilojoules, released when 45.8 g of steam (water) condenses at its boiling (condensation) point?

Answer: 45.8 g 45.8 g steam × 2260 J 1 g steam = 104 kJ

Three SFs Exact Exact Three SFs
The chapter sections to review are shown in parentheses at the end of each question.

10.59 State the number of valence electrons, bonding pairs, and lone pairs in each of the following Lewis structures: (10.1)

a. H: H
b. H: Br

10.60 State the number of valence electrons, bonding pairs, and lone pairs in each of the following Lewis structures: (10.1)

a. H: O: H
b. H: N: H

10.61 Match each of the Lewis structures (a to c) with the correct diagram (1 to 3) of its shape, and name the shape; indicate if each molecule is polar or nonpolar. Assume X and Y are nonmetals and all bonds are polar covalent. (10.1, 10.3, 10.5)

10.62 Match each of the formulas (a to c) with the correct diagram (1 to 3) of its shape, and name the shape; indicate if each molecule is polar or nonpolar. (10.1, 10.3, 10.5)

10.63 Consider the following bonds: C — C, C — H, C — F, C — Li, and C — Mg. (10.4)

a. Which bonds are polar covalent?
b. Which bonds are nonpolar covalent?
c. Which bonds are ionic?
d. Arrange the covalent bonds in order of decreasing polarity.

10.64 Consider the following bonds: Li — H, Na — H, C — H, O — H, and Cl — H. (10.4)

a. Which bonds are polar covalent?
b. Which bonds are nonpolar covalent?
c. Which bonds are ionic?
d. Arrange the covalent bonds in order of decreasing polarity.

10.65 Identify the major intermolecular forces between each of the following atoms or molecules: (10.6)

a. PCl₃
b. CH₃CH₂OH
c. CH₃CH₂NH₃
d. O₂

d. Arrange the covalent bonds in order of decreasing polarity.

10.66 Identify the major intermolecular forces between each of the following atoms or molecules: (10.6)

a. CS₂
b. HCl
c. CCl₄
d. Kr

d. Arrange the covalent bonds in order of decreasing polarity.

10.67 Use your knowledge of changes of state to explain the following: (10.7)

a. How does perspiration during heavy exercise cool the body?
b. Why do towels dry more quickly on a hot summer day than on a cold winter day?
c. Why do wet clothes stay wet in a plastic bag?

10.68 Use your knowledge of changes of state to explain the following: (10.7)

a. Why is a spray that evaporates quickly, such as ethyl chloride, used to numb a sports injury during a game?
b. Why does water in a wide, flat, shallow dish evaporate more quickly than the same amount of water in a tall, narrow vase?
c. Why does a sandwich on a plate dry out faster than a sandwich in plastic wrap?

10.69 Draw a heating curve for a sample of ice that is heated from −20 °C to 150 °C. Indicate the segment of the graph that corresponds to each of the following: (10.7)
a. solid
b. melting
c. liquid
d. boiling
e. gas

10.70 Draw a cooling curve for a sample of steam that cools from 110 °C to −10 °C. Indicate the segment of the graph that corresponds to each of the following: (10.7)
a. solid
b. freezing
c. liquid
d. condensing
e. gas
10.71 The following is a heating curve for chloroform, a solvent for fats, oils, and waxes: (10.5)

![Heating Curve for Chloroform](image)

a. What is the approximate melting point of chloroform?
b. What is the approximate boiling point of chloroform?
c. On the heating curve, identify the segments A, B, C, D, and E as solid, liquid, gas, melting, or boiling.
d. At the following temperatures, is chloroform a solid, liquid, or gas? -80 °C; -40 °C; 25 °C; 80 °C

10.72 Associate the contents of the beakers (1 to 5) with segments (A to E) on the heating curve for water. (10.7)

![Heating Curve for Water](image)

10.73 Determine the total number of valence electrons in each of the following: (10.1)
- a. HNO₂
- b. CH₃CHO
- c. PH₄⁺
- d. SO₃²⁻

10.74 Determine the total number of valence electrons in each of the following: (10.1)
- a. COCl₂
- b. N₂O
- c. BrO₂⁻
- d. SeCl₂

10.75 Draw the Lewis structure for each of the following: (10.1)
- a. BF₄⁻
- b. Cl₂O
- c. H₂NOH (N is the central atom)
- d. H₂CCl₂

10.76 Draw the Lewis structure for each of the following: (10.1)
- a. H₃COCH₃ (the atoms are in the order C O C)
- b. HNO₂ (the atoms are in the order HONO)
- c. IO₃⁻
- d. BrO⁻

10.77 Draw resonance structures for each of the following: (10.2)
- a. N₃⁻
- b. NO₂⁺
- c. HCO₂⁻

10.78 Draw resonance structures for each of the following: (10.2)
- a. NO₃⁻
- b. CO₂⁺
- c. SCN⁻

10.79 Use the periodic table to arrange the following atoms in order of increasing electronegativity: (10.4)
- a. N, P, As
- b. Ba, Be, B, Bi
- c. Na, K, Br, Al

10.80 Use the periodic table to arrange the following atoms in order of increasing electronegativity: (10.4)
- a. Tl, Ga, P
- b. O, Se, Ga, K
- c. Po, C, N, O

10.81 Select the more polar bond in each of the following pairs: (10.4)
- a. H—I or H—Cl
- b. P—F or P—Cl
- c. Si—H or C—H
- d. Si—O or C—O
- e. N—Cl or N—F

10.82 Select the more polar bond in each of the following pairs: (10.4)
- a. H—B or H—C
- b. O—F or O—Br
- c. Cl—O or Br—O
- d. H—Al or H—B
- e. N—C or P—C

10.83 Show the dipole arrow for each of the following bonds: (10.4)
- a. C—P
- b. O—F
- c. O—Si
- d. Br—H
- e. Cl—I

10.84 Show the dipole arrow for each of the following bonds: (10.4)
- a. N—P
- b. Si—F
- c. O—Se
- d. Ga—O
- e. P—Br

10.85 Calculate the electronegativity difference and classify each of the following bonds as nonpolar covalent, polar covalent, or ionic: (10.4)
- a. Si—Cl
- b. C—C
- c. Na—Cl
- d. C—H
- e. F—F
10.86 Calculate the electronegativity difference and classify each of the following bonds as nonpolar covalent, polar covalent, or ionic: (10.4)
   a. C—N  
   b. Cl—Cl  
   c. K—Br  
   d. H—H  
   e. N—F

10.87 For each of the following, draw the Lewis structure and determine the shape: (10.1, 10.3)
   a. NF3  
   b. SiBr4  
   c. CSe2  
   d. SO2

10.88 For each of the following, draw the Lewis structure and determine the shape: (10.1, 10.3)
   a. PCl4+  
   b. O2  
   c. COCl2 (C is the central atom)  
   d. HCCH

10.89 Use the Lewis structure to determine the shape for each of the following molecules or polyatomic ions: (10.1, 10.3)
   a. BrO2-  
   b. H2O  
   c. CBr4  
   d. PO3  

10.90 Predict the shape and polarity of each of the following molecules, which have polar covalent bonds: (10.3, 10.5)
   a. A central atom with three identical bonded atoms and one lone pair.  
   b. A central atom with two bonded atoms and two lone pairs.

10.91 Predict the shape and polarity of each of the following molecules, which have polar covalent bonds: (10.3, 10.5)
   a. A central atom with four identical bonded atoms and no lone pairs.  
   b. A central atom with four bonded atoms that are not identical and no lone pairs.

10.92 Classification of molecules or ions:
   a. HBBr  
   b. SiO2  
   c. NCl3  
   d. CH4Cl  
   e. Ni3  
   f. H2O

10.93 For each of the following molecules as polar or nonpolar:
   a. HBr  
   b. SiO2  
   c. NCl3  
   d. CH4Cl  
   e. Ni3  
   f. H2O

10.94 Indicate the major types of intermolecular forces—
   1. ionic bonds, 2. dipole–dipole attractions, 3. hydrogen bonds, 4. dispersion forces—that occur between particles of the following: (10.6)
   a. SiHCl3  
   b. BH3  
   c. Ar  
   d. PH3  
   e. BF3  
   f. LiOH

10.95 When it rains or snows, the air temperature seems warmer. Explain. (10.7)

10.96 Water is sprayed on the ground of an orchard when temperatures are near freezing to keep the fruit from freezing. Explain. (10.7)

10.97 Using Figure 10.6, calculate the grams of ice that will melt at 0°C if 1540 J is absorbed. (10.7)

10.98 Using Figure 10.6, calculate the grams of ethanol that will vaporize at its boiling point if 4620 J is absorbed. (10.7)

10.99 Using Figure 10.6, calculate the grams of acetic acid that will freeze at its freezing point if 5.25 kJ is removed. (10.7)

10.100 Using Figure 10.6, calculate the grams of benzene that will condense at its boiling point if 8.46 kJ is removed. (10.7)

10.101 Identify the errors in each of the following Lewis structures and draw the correct formula: (10.1)
   a. Cl  
   b. H—C—H  
   c. H—N—O—H

10.102 Complete the Lewis structure for each of the following: (10.1)
   a. H—N—C—H  
   b. Cl—C—N—H  
   c. Cl—O—C—H

10.103 Complete the Lewis structure for each of the following: (10.1)
   a. H—N—C—H  
   b. Cl—C—N—H  
   c. H—N—O—H  
   d. Cl—O—C—H

10.104 Predict the shape of each of the following molecules or ions: (10.3)
   a. HNO  
   b. PCl4+  
   c. BCl3

10.105 Identify the errors in each of the following Lewis structures and draw the correct formula: (10.1)
   a. Cl—O—Cl  
   b. H—C—H  
   c. H—N—O—H

10.106 Predict the shape of each of the following molecules or ions: (10.3, 10.5)
   a. BF3  
   b. HOCl (O is the central atom)  
   c. N(CH3)3 (N is the central atom)
10.107 The melting point of dibromomethane is −53 °C and its boiling point is 97 °C. Sketch a heating curve for dibromomethane from −100 °C to 120 °C. (10.7)
   a. What is the state of dibromomethane at −75 °C?
   b. What happens on the curve at −53 °C?
   c. What is the state of dibromomethane at −18 °C?
   d. What is the state of dibromomethane at 110 °C?
   e. At what temperature will both solid and liquid be present?

10.108 The melting point of benzene is 5.5 °C and its boiling point is 80.1 °C. Draw a heating curve for benzene from 0 °C to 100 °C. (10.7)
   a. What is the state of benzene at 15 °C?
   b. What happens on the curve at 5.5 °C?
   c. What is the state of benzene at 63 °C?
   d. What is the state of benzene at 98 °C?
   e. At what temperature will both liquid and gas be present?

10.109 A 45.0-g piece of ice at 0.0 °C is added to a sample of water at 8.0 °C. All of the ice melts, and the temperature of the water decreases to 0.0 °C. How many grams of water were in the sample? (10.7)

10.110 An ice cube at 0 °C with a mass of 115 g is added to water in a beaker that has a temperature of 64.0 °C. If the final temperature of the mixture is 24.0 °C, what was the initial mass of the warm water? (10.7)

10.111 A pot holds 250 g of water. If the water initially has a temperature of 20 °C, how many kilojoules of heat is required to boil the water at 100 °C? (10.7)

10.112 A 3.0-kg block of lead is taken from a furnace at 10.112 °C and placed on a large block of ice at 0.0 °C. If all the heat given up by the lead is used to melt ice, how much ice is melted when the temperature of the lead drops to 0.0 °C? (10.7)

**ANSWERS**

Answers to Selected Questions and Problems

10.1 a. 8 valence electrons 
   b. 14 valence electrons 
   c. 32 valence electrons

10.3 a. HF (8 e\(^{-}\)) or H\(\overset{\cdot}{F}\);
   b. SF\(_2\) (20 e\(^{-}\)) or F\(\overset{\cdot}{S}\)\(\overset{\cdot}{F}\); or F\(\overset{\cdot}{S}\)\(\overset{\cdot}{F}\);
   c. NBr\(_3\) (26 e\(^{-}\)) or Br\(\overset{\cdot}{N}\)\(\overset{\cdot}{Br}\); or Br\(\overset{\cdot}{N}\)\(\overset{\cdot}{Br}\);
   d. BH\(_4\)\(^-\) (8 e\(^{-}\)) or H\(\overset{\cdot}{B}\)\(\overset{\cdot}{H}\); or H\(\overset{\cdot}{B}\)\(\overset{\cdot}{H}\);
   e. CH\(_3\)OH (14 e\(^{-}\)) or H\(\overset{\cdot}{C}\)\(\overset{\cdot}{O}\)\(\overset{\cdot}{H}\); or H\(\overset{\cdot}{C}\)\(\overset{\cdot}{O}\)\(\overset{\cdot}{H}\);
   f. N\(_2\)H\(_4\) (14 e\(^{-}\)) or H\(\overset{\cdot}{N}\)\(\overset{\cdot}{N}\)\(\overset{\cdot}{H}\); or H\(\overset{\cdot}{N}\)\(\overset{\cdot}{N}\)\(\overset{\cdot}{H}\).

10.5 If complete octets cannot be formed by using all the valence electrons, it is necessary to draw multiple bonds.

10.7 a. CO (10 e\(^{-}\)) or C\(\overset{\cdot}{O}\); or C\(\overset{\cdot}{\equiv}\)O;
   b. CN\(^-\) (10 e\(^{-}\)) or C\(\overset{\cdot}{\equiv}\)N\(^-\);
   c. H\(_2\)CO (12 e\(^{-}\)) or H\(\overset{\cdot}{C}\)\(\overset{\cdot}{O}\); or H\(\overset{\cdot}{C}\)\(\overset{\cdot}{O}\).

10.9 Resonance occurs when we can draw two or more electron-dot formulas for the same molecule or ion.

10.10 a. ClNO\(_2\) (24 e\(^{-}\)) or Cl\(\overset{\cdot}{N}\)\(\overset{\cdot}{O}\); or Cl\(\overset{\cdot}{N}\)\(\overset{\cdot}{O}\);
   b. OCN\(^-\) (16 e\(^{-}\)) or O\(\overset{\cdot}{C}\)\(\overset{\cdot}{N}\); or O\(\overset{\cdot}{C}\)\(\overset{\cdot}{N}\).

10.13 a. 6, tetrahedral
   b. 5, trigonal pyramidal
   c. 3, trigonal planar

10.15 a. three
   b. trigonal planar
   c. three
   d. trigonal planar

10.17 In CF\(_3\), the central atom C has four bonded atoms and no lone pairs of electrons, which gives it a tetrahedral shape. In NF\(_3\), the central atom N has three bonded atoms and one lone pair of electrons, which gives NF\(_3\) a trigonal pyramidal shape.

10.19 a. trigonal planar
   b. bent (109°)
   c. linear
   d. tetrahedral

10.21 a. AlH\(_4\)\(^-\) (8 e\(^{-}\)) or H\(\overset{\cdot}{A}\)\(\overset{\cdot}{l}\)\(\overset{\cdot}{H}\); or H\(\overset{\cdot}{A}\)\(\overset{\cdot}{l}\)\(\overset{\cdot}{H}\);
   b. SO\(_4\)^{2-} (32 e\(^{-}\)) or O\(\overset{\cdot}{S}\)\(\overset{\cdot}{O}\); or O\(\overset{\cdot}{S}\)\(\overset{\cdot}{O}\);
   c. NH\(_4\)\(^+\) (8 e\(^{-}\)) or H\(\overset{\cdot}{N}\)\(\overset{\cdot}{H}\); or H\(\overset{\cdot}{N}\)\(\overset{\cdot}{H}\);
   d. NO\(_2\)^{2+} (16 e\(^{-}\)) or O\(\overset{\cdot}{N}\)\(\overset{\cdot}{O}\); or O\(\overset{\cdot}{N}\)\(\overset{\cdot}{O}\).

10.23 a. increases
   b. decreases
   c. decreases

10.25 a. between 0.0 and 0.4

10.27 a. K, Na, Li
   b. Na, P, Cl
   c. Ca, Se, O
10.29 a. polar covalent  b. ionic
c. polar covalent  d. nonpolar covalent
e. polar covalent  f. nonpolar covalent

10.31 a. \( \text{N}^-\text{F} \)  b. \( \text{Si}^+\text{Br}^- \)
c. \( \text{C}^-\text{O}^- \)  d. \( \text{P}^+\text{Br}^- \)
e. \( \text{N}^-\text{P}^+ \)

10.33 Electrons are shared equally between two identical atoms and unequally between nonidentical atoms.

10.35 a. nonpolar  b. polar
c. polar  d. nonpolar

10.37 In the molecule \( \text{CO}_2 \), the two \( \text{C}^-\text{O}^- \) dipoles cancel; in \( \text{CO} \), there is only one dipole.

10.39 a. dipole–dipole attractions  b. ionic bonds
c. dipole–dipole attractions  d. dispersion forces
d. dispersion forces

10.41 a. hydrogen bonds  b. dipole–dipole attractions
c. dispersion forces  d. dispersion forces

10.43 a. HF; hydrogen bonds are stronger than the dipole–dipole attractions in \( \text{HBr} \).
b. \( \text{NaF} \); ionic bonds are stronger than the hydrogen bonds in HF.
c. MgBr\(_2\); ionic bonds are stronger than the dipole–dipole attractions in \( \text{PBBr}_3 \).
d. \( \text{CH}_3\text{OH} \); hydrogen bonds are stronger than the dispersion forces in \( \text{CH}_4 \).

10.45 a. 21 700 J  b. 2180 J
c. 22.3 kJ  d. 16.7 kJ

10.47 a. 22 600 J  b. 42.1 kJ
c. 3100 J  d. 242 kJ

10.49 a. 4800 J  b. 30 300 J
c. 40.2 kJ  d. 72.2 kJ

10.51 119.5 kJ

10.53 3120 J

10.55 a. 0.0, nonpolar covalent  b. 0.4, nonpolar covalent
c. 1.0, polar covalent

10.57 a.
\[
\begin{align*}
\text{[O=C=O]}^2- & \leftrightarrow \text{[O=C=O]}^2- \\
\text{[O=O]}^- & \leftrightarrow \text{[O=O]}^- \\
\text{[O=O]}^- & \leftrightarrow \text{[O=O]}^- 
\end{align*}
\]

b. Carbonate ion is trigonal planar.

10.59 a. two valence electrons, one bonding pair, no lone pairs
b. eight valence electrons, one bonding pair, three lone pairs
c. 14 valence electrons, one bonding pair, six lone pairs

10.61 a. 2, trigonal pyramidal, polar  b. 1, bent (109°), polar
c. 3, tetrahedral, nonpolar

c. None  d. C—F, C—H, C—C

c. Dipole–dipole attractions  d. Dispersion forces

10.67 a. The heat from the skin is used to evaporate the water (perspiration). Therefore, the skin is cooled.
b. On a hot day, there are more molecules with sufficient energy to become water vapor.
c. In a closed bag, some molecules evaporate, but they cannot escape and will condense back to liquid; the clothes will not dry.

10.69

10.71 a. about −60 °C  b. about 60 °C
c. The diagonal line A represents the solid state as temperature increases. The horizontal line B represents the change from solid to liquid or melting of the substance. The diagonal line C represents the liquid state as temperature increases. The horizontal line D represents the change from liquid to gas or boiling of the liquid. The diagonal line E represents the gas state as temperature increases.
d. At −80 °C, solid; at −40 °C, liquid; at 25 °C, liquid; 80 °C, gas

10.73 a. \( 1 + 5 + 2(6) = 18 \) valence electrons
b. \( 2(4) + 4(1) + 6 = 18 \) valence electrons
c. \( 5 + 4(1) - 1 = 8 \) valence electrons
b. \( 6 + 3(6) + 2 = 26 \) valence electrons

10.75 a. BF\(_4^-\) (32 e\(^-\))  \( \text{[F-B-F]}^- \) or \( \text{[F-B-F]}^- \)
b. Cl\(_2\)O (20 e\(^-\))  \( \text{[Cl}^-\text{O}^-\text{Cl}^-\text{]} \) or \( \text{[Cl}^-\text{O}^-\text{Cl}^-\text{]} \)
c. H\(_2\)NOH (14 e\(^-\))  H\(_2\)N\text{[O=O]}H \ or \ H\(_2\)N\text{[O=O]}H

d. H\(_2\)CCCl\(_2\) (24 e\(^-\))  H\(_2\)C\text{[Cl]}\(_2\) \ or \ H\(_2\)C\text{[Cl]}\(_2\)

e. H\(_2\)NOH (16 e\(^-\))  \( \text{[N=O]^-\text{N=O}^-\text{]} } \) or \( \text{[N=O]^-\text{N=O}^-\text{]} } 

10.77 a. (16 e\(^-\))  \( \text{[N=N=N]}^- \) or \( \text{[N=N=N]}^- \)
b. (16 e\(^-\))  \( \text{[O=O]}^- \) or \( \text{[O=O]}^- \)
c. (18 e\(^-\))  \( \text{[O=O]}^- \) or \( \text{[O=O]}^- \)
10.79 a. As, P, N  b. Ba, Be, Bi, B  c. K, Na, Al, Br


10.85 a. polar covalent  b. nonpolar covalent  c. ionic  d. nonpolar covalent  e. nonpolar covalent

10.87 a. NF₃ (26 e⁻)  :F—N—F;  trigonal pyramidal  b. SiBr₄ (32 e⁻)  :Br—Si—Br;  tetrahedral  c. CSe₂ (16 e⁻)  :Se=C=Se;  linear  d. SO₂ (18 e⁻)  :O=S—O;  bent (120°)

10.89 a. bent (109°)  b. bent (109°)  c. tetrahedral  d. trigonal pyramidal

10.91 a. trigonal pyramidal, polar  b. bent (109°), polar

10.93 a. polar  b. nonpolar  c. nonpolar  d. polar  e. polar  f. polar

10.95 a. (2) dipole–dipole attractions  b. (4) dispersion forces  c. (4) dispersion forces  d. (2) dipole–dipole attractions  e. (4) dispersion forces  f. (1) ionic bonds

10.97 When water vapor condenses or liquid water freezes, heat is released, which warms the air.

10.99 4.61 g of water

10.101 27.3 g of acetic acid

10.103 a. (18 e⁻)  H—N—C—H  b. (22 e⁻)  :Cl—C≡N::H  c. (12 e⁻)  H—N=N—H  d. (30 e⁻)  :Cl—C—O—C—H

10.105 a. linear  b. tetrahedral  c. trigonal planar

10.109 450 g of water

10.111 648.6 kJ is required.
CI.13 In an experiment, the mass of a piece of copper is determined to be 8.56 g. Then the copper is reacted with sufficient oxygen gas to produce solid copper(II) oxide.

(7.1, 8.1, 8.2, 8.3, 9.1, 9.2, 9.3, 9.5)

a. How many copper atoms are in the sample?
b. Write the balanced chemical equation for the reaction.
c. Classify the type of reaction.
d. How many grams of O₂ are required to completely react with the Cu?
e. How many grams of CuO will result from the reaction of 8.56 g of Cu and 3.72 g of oxygen?
f. How many grams of CuO will result in part e, if the yield for the reaction is 85.0%?

CI.14 One of the components in gasoline is octane, C₈H₁₈, which has a density of 0.803 g/cm³ and ΔH = −510 kJ/mol. Suppose a hybrid car has a fuel tank with a capacity of 11.9 gal and has a gas mileage of 45 mi/gal.

(1.10, 6.7, 8.3, 8.5, 9.6)

a. Write a balanced chemical equation for the complete combustion of octane including the heat of reaction.
b. What is the energy, in kilojoules, produced from one tank of fuel assuming it is all octane?
c. How many molecules of C₈H₁₈ are present in one tank of fuel assuming it is all octane?
d. If this hybrid car is driven 24 500 miles in one year, how many kilograms of carbon dioxide will be produced from the combustion of the fuel assuming it is all octane?

CI.15 When clothes have stains, bleach may be added to the wash to react with the soil and make the stains colorless. The bleach solution is prepared by bubbling chlorine gas into a solution of sodium hydroxide to produce a solution of sodium hypochlorite, sodium chloride, and water. One brand of bleach contains 5.25% sodium hypochlorite by mass (active ingredient) with a density of 1.08 g/mL.

(6.4, 7.1, 7.2, 8.2, 9.3, 9.4, 9.5, 10.1)

a. What is the formula and molar mass of sodium hypochlorite?
b. Draw the Lewis structure for the hypochlorite ion.
c. How many hypochlorite ions are present in 1.00 gal of bleach solution?
d. Write the balanced chemical equation for the preparation of bleach.
e. How many grams of NaOH are required to produce the sodium hypochlorite for 1.00 gal of bleach?
f. If 165 g of Cl₂ is passed through a solution containing 275 g of NaOH and 162 g of sodium hypochlorite is produced, what is the percent yield of sodium hypochlorite for the reaction?

CI.16 Ethanol, C₂H₆O, is obtained from renewable crops such as corn, which use the Sun as their source of energy. In the United States, automobiles can now use a fuel known as E85 that contains 85.0% ethanol and 15.0% gasoline by volume. Ethanol has a melting point of −115 °C, a boiling point of 78 °C, a heat of fusion of 109 J/g, and a heat of vaporization of 841 J/g. Liquid ethanol has a density of 0.796 g/mL and a specific heat of 2.46 J/g °C.

(8.3, 8.5, 9.2, 9.3, 10.4, 10.5)

E85 fuel contains 85% ethanol.

a. Draw a heating curve for ethanol from −150 °C to 100 °C.

b. When 20.0 g of ethanol at −62 °C is heated and completely vaporized at 78 °C, how much energy, in kilojoules, is required?
c. If a 15.0-gal gas tank is filled with E85, how many liters of ethanol are in the gas tank?
d. Write the balanced chemical equation for the complete combustion of ethanol.

e. How many kilograms of CO₂ are produced by the complete combustion of the ethanol in a full 15.0-gal gas tank?

f. What would be the strongest intermolecular force between liquid ethanol molecules?

Applications

CI.17 Chloral hydrate, a sedative and hypnotic, was the first drug used to treat insomnia. Chloral hydrate has a melting point of 57 °C. At its boiling point of 98 °C, it breaks down to chloral and water. (7.4, 7.5, 8.4, 10.1)

\[
\begin{align*}
\text{Cl} & \quad \text{O} \\
\text{Cl} & \quad \text{C} \quad \text{C} \quad \text{O} \quad \text{H} & \quad \text{Cl} & \quad \text{O} \\
\text{Cl} & \quad \text{C} & \quad \text{C} & \quad \text{H} + \text{H}_2\text{O} \\
\text{Cl} & \quad \text{H} & \quad \text{Cl} & \quad \text{Cl}
\end{align*}
\]

Chloral hydrate \quad \text{Chloral}

a. Draw the Lewis structures for chloral hydrate and chloral.
b. What are the empirical formulas of chloral hydrate and chloral?
c. What is the mass percent of Cl in chloral hydrate?

CI.18 Ethylene glycol, C₂H₆O₂, used as a coolant and antifreeze, has a density of 1.11 g/mL. As a sweet-tasting liquid, it can be appealing to pets and small children, but it can be toxic, with an LD₅₀ of 4700 mg/kg. Its accidental ingestion can cause kidney damage and difficulty with breathing. In the body, ethylene glycol is converted to another toxic substance, oxalic acid, H₂C₂O₄. (2.8, 7.4, 7.5, 8.2, 8.4, 10.1, 10.3, 10.4)

a. What are the empirical formulas of ethylene glycol and oxalic acid?
b. If ethylene glycol has a C—C single bond with two H atoms attached to each C atom, what is its Lewis structure?
c. Which bonds in ethylene glycol are polar covalent and which are nonpolar covalent?
d. How many milliliters of ethylene glycol could be toxic for an 11.0-lb cat?
e. What would be the strongest intermolecular force in ethylene glycol?

f. If oxalic acid has two carbon atoms attached by a C—C single bond with each carbon also attached to two oxygen atoms, what is its Lewis structure?
g. Write the balanced chemical equation for the reaction of ethylene glycol and oxygen (O₂) to give oxalic acid and water.

CI.19 Acetone (propanone), a clear liquid solvent with an acrid odor, is used to remove nail polish, paints, and resins. It has a low boiling point and is highly flammable. Acetone has a density of 0.786 g/mL and a heat of combustion of −1790 kJ/mol. (2.8, 7.2, 8.3, 8.4, 9.3, 9.6)

\[
\begin{align*}
\text{Cl} & \quad \text{O} \\
\text{Cl} & \quad \text{C} \quad \text{C} \quad \text{H} \\
\text{Cl} & \quad \text{H} & \quad \text{Cl}
\end{align*}
\]

Acetone has carbon atoms (black), hydrogen atoms (white), and an oxygen atom (red).

a. Draw the Lewis structure for acetone.
b. What are the molecular formula and molar mass of acetone?
c. Write a balanced chemical equation for the complete combustion of acetone, including the heat of reaction.
d. Is the combustion of acetone an endothermic or exothermic reaction?
e. How much heat, in kilojoules, is released if 2.58 g of acetone reacts completely with oxygen?
f. How many grams of oxygen gas are needed to react with 15.0 mL of acetone?

CI.20 The compound dihydroxyacetone (DHA) is used in “sunless” tanning lotions, which darken the skin by reacting with the amino acids in the outer surface of the skin. A typical drugstore lotion contains 4.0% (mass/volume) DHA. (2.8, 7.2, 8.4)

a. Draw the Lewis structure for DHA.
b. DHA has C—C, C—H, C—O, and O—H bonds. Which of these bonds are polar covalent and which are nonpolar covalent?
c. What are the molecular formula and molar mass of DHA?
d. A bottle of sunless tanning lotion contains 177 mL of lotion. How many milligrams of DHA are in a bottle?
ANSWERS

CI.13  a. $8.11 \times 10^{22}$ atoms of copper
b. $2\text{Cu}(s) + \text{O}_2(g) \rightarrow 2\text{CuO}(s)$
c. combination reaction
d. 2.16 g of O₂
e. 10.7 g of CuO
f. 9.10 g of CuO

CI.15  a. NaOCl, 74.44 g/mol
b. $\left(14e^-\right)$ \[\begin{array}{c} \text{Cl} \\ \text{O} \end{array} \] 
c. $1.74 \times 10^{24}$ OCl⁻ ions
d. $2\text{NaOH}(aq) + \text{Cl}_2(g) \rightarrow \text{NaOCl}(aq) + \text{NaCl}(aq) + \text{H}_2\text{O}(l)$
e. 231 g of NaOH
f. 93.6%

CI.17  a. $\left(44e^-\right)$ \[\begin{array}{c} \text{Cl} \\ \text{O} \end{array} \] 
b. \[\begin{array}{c} \text{Cl} \\ \text{C} \\ \text{H} \end{array} \] 
c. \[\begin{array}{c} \text{O} \\ \text{C} \\ \text{O} \end{array} \]

CI.19  a. \[\begin{array}{c} \text{CH}_3 \longrightarrow \text{C} \longrightarrow \text{CH}_3 \end{array} \]
b. C₃H₆O; 58.08 g/mol
c. \[\begin{array}{c} \text{C}_3\text{H}_6\text{O}(g) + 4\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 3\text{H}_2\text{O}(g) + 1790 \text{kJ} \]
d. exothermic
e. 79.5 kJ
f. 26.0 g of O₂
AFTER SOCCER PRACTICE, Whitney complained that she was having difficulty breathing. Her father took her to the emergency room where she was seen by Sam, a respiratory therapist, who listened to Whitney's chest and then tested her breathing capacity using a spirometer. Based on her limited breathing capacity and the wheezing noise in her chest, Whitney was diagnosed as having asthma.

Sam gave Whitney a nebulizer containing a bronchodilator that opens the airways and allows more air to go into the lungs. During the breathing treatment, he measured the amount of oxygen (O₂) in her blood and explained to Whitney and her father that air is a mixture of gases containing 78% nitrogen (N₂) gas and 21% O₂ gas. Because Whitney had difficulty obtaining sufficient oxygen, Sam gave her supplemental oxygen through an oxygen mask. Within a short period of time, Whitney's breathing returned to normal. The therapist then explained that the lungs work according to Boyle's law: The volume of the lungs increases upon inhalation, and the pressure decreases to allow air to flow in. However, during an asthma attack, the airways become restricted, and it becomes more difficult to expand the volume of the lungs.

CAREER

Respiratory Therapist

Respiratory therapists assess and treat a range of patients, including premature infants whose lungs have not developed and asthmatics or patients with emphysema or cystic fibrosis. In assessing patients, they perform a variety of diagnostic tests including breathing capacity, concentrations of oxygen and carbon dioxide in a patient's blood, as well as blood pH. In order to treat patients, therapists provide oxygen or aerosol medications to the patient, as well as chest physiotherapy to remove mucus from their lungs. Respiratory therapists also educate patients on how to correctly use their inhalers.
11.1 Properties of Gases

**LEARNING GOAL.** Describe the kinetic molecular theory of gases and the units of measurement used for gases.

We all live at the bottom of a sea of gases called the atmosphere. The most important of these gases is oxygen, which constitutes about 21% of the atmosphere. Without oxygen, life on this planet would be impossible: Oxygen is vital to all life processes of plants and animals. Ozone (O₃), formed in the upper atmosphere by the interaction of oxygen with ultraviolet light, absorbs some of the harmful radiation before it can strike Earth’s surface. The other gases in the atmosphere include nitrogen (78%), argon, carbon dioxide (CO₂), and water vapor. Carbon dioxide gas, a product of combustion and metabolism, is used by plants in photosynthesis, which produces the oxygen that is essential for humans and animals.

The behavior of gases is quite different from that of liquids and solids. Gas particles are far apart, whereas particles of both liquids and solids are held close together. A gas has no definite shape or volume and will completely fill any container. Because there are great distances between gas particles, a gas is less dense than a solid or liquid, and easy to compress. A model for the behavior of a gas, called the **kinetic molecular theory of gases**, helps us understand gas behavior.

**Kinetic Molecular Theory of Gases**

1. **A gas consists of small particles (atoms or molecules) that move randomly with high velocities.** Gas molecules moving in random directions at high speeds cause a gas to fill the entire volume of a container.
2. **The attractive forces between the particles of a gas are usually very small.** Gas particles are far apart and fill a container of any size and shape.
3. **The actual volume occupied by gas molecules is extremely small compared to the volume that the gas occupies.** The volume of the gas is considered equal to the volume of the container. Most of the volume of a gas is empty space, which allows gases to be easily compressed.
4. **Gas particles are in constant motion, moving rapidly in straight paths.** When gas particles collide, they rebound and travel in new directions. Every time they hit the walls of the container, they exert pressure. An increase in the number or force of collisions against the walls of the container causes an increase in the pressure of the gas.
5. **The average kinetic energy of gas molecules is proportional to the Kelvin temperature.** Gas particles move faster as the temperature increases. At higher temperatures, gas particles hit the walls of the container more often and with more force, producing higher pressures.
The kinetic molecular theory helps explain some of the characteristics of gases. For example, you can smell perfume when a bottle is opened on the other side of a room because its particles move rapidly in all directions. At room temperature, the molecules in the air are moving at about 450 m/s, which is 1000 mi/h. They move faster at higher temperatures and more slowly at lower temperatures. Sometimes tires and gas-filled containers explode when temperatures are too high. From the kinetic molecular theory, you know that gas particles move faster when heated, hit the walls of a container with more force, and cause a buildup of pressure inside a container.

When we talk about a gas, we describe it in terms of four properties: pressure, volume, temperature, and the amount of gas.

### Pressure ($P$)

Gas particles are extremely small and move rapidly. When they hit the walls of a container, they exert a pressure (see FIGURE 11.1). If we heat the container, the molecules move faster and smash into the walls of the container more often and with increased force, thus increasing the pressure. The gas particles in the air, mostly oxygen and nitrogen, exert a pressure on us called atmospheric pressure (see FIGURE 11.2). As you go to higher altitudes, the atmospheric pressure is less because there are fewer particles in the air. The most common units used to measure gas pressure are the atmosphere (atm) and millimeters of mercury (mmHg). On the TV weather report, you may hear or see the atmospheric pressure given in inches of mercury, or in kilopascals in countries other than the United States. In a hospital, the unit torr or pounds per square inch (psi) may be used.

**FIGURE 11.1** Gas particles move in straight lines within a container. The gas particles exert pressure when they collide with the walls of the container.

**Why does heating the container increase the pressure of the gas within it?**

**FIGURE 11.2** A column of air extending from the upper atmosphere to the surface of Earth produces a pressure on each of us of about 1 atm. While there is a lot of pressure on the body, it is balanced by the pressure inside the body.

**Why is there less atmospheric pressure at higher altitudes?**
Volume (V)
The volume of gas equals the size of the container in which the gas is placed. When you inflate a tire or a basketball, you are adding more gas particles. The increase in the number of particles hitting the walls of the tire or basketball increases the volume. Sometimes, on a cold morning, a tire looks flat. The volume of the tire has decreased because a lower temperature decreases the speed of the molecules, which in turn reduces the force of their impacts on the walls of the tire. The most common units for volume measurement are liters (L) and milliliters (mL).

Temperature (T)
The temperature of a gas is related to the kinetic energy of its particles. For example, if we have a gas at 200 K and heat it to a temperature of 400 K, the gas particles will have twice the kinetic energy that they did at 200 K. This also means that the gas at 400 K exerts twice the pressure of the gas at 200 K. Although you measure gas temperature using a Celsius thermometer, all comparisons of gas behavior and all calculations related to temperature must use the Kelvin temperature scale. No one has quite created the conditions for absolute zero (0 K), but we predict that the particles will have zero kinetic energy and exert zero pressure at absolute zero.

Amount of Gas (n)
When you add air to a bicycle tire, you increase the amount of gas, which results in a higher pressure in the tire. Usually, we measure the amount of gas by its mass, in grams. In gas law calculations, we need to change the grams of gas to moles. A summary of the four properties of a gas is given in Table 11.1.

### Table 11.1 Properties That Describe a Gas

<table>
<thead>
<tr>
<th>Property</th>
<th>Description</th>
<th>Units of Measurement</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pressure (P)</td>
<td>The force exerted by a gas against the walls of the container</td>
<td>atmosphere (atm); millimeter of mercury (mmHg); torr; pascal (Pa)</td>
</tr>
<tr>
<td>Volume (V)</td>
<td>The space occupied by a gas</td>
<td>liter (L); milliliter (mL)</td>
</tr>
<tr>
<td>Temperature (T)</td>
<td>The determining factor of the kinetic energy and rate of motion of gas particles</td>
<td>degree Celsius (°C); kelvin (K) is required in calculations</td>
</tr>
<tr>
<td>Amount (n)</td>
<td>The quantity of gas present in a container</td>
<td>gram (g); mole (n) is required in calculations</td>
</tr>
</tbody>
</table>

### Sample Problem 11.1 Properties of Gases
Identify the property of a gas that is described by each of the following:

a. increases the kinetic energy of gas particles
b. the force of the gas particles hitting the walls of the container
c. the space that is occupied by a gas

#### Try It First

**Solution**

a. temperature  

b. pressure  

c. volume  

### Study Check 11.1

When helium is added to a balloon, the number of grams of helium increases. What property of a gas is described?

#### Answer

The mass, in grams, gives the amount of gas.
Your blood pressure is one of the vital signs a doctor or nurse checks during a physical examination. It actually consists of two separate measurements. Acting like a pump, the heart contracts to create the pressure that pushes blood through the circulatory system. During contraction, the blood pressure is at its highest; this is your systolic pressure. When the heart muscles relax, the blood pressure falls; this is your diastolic pressure. The normal range for systolic pressure is 100 to 200 mmHg. For diastolic pressure, it is 60 to 80 mmHg. These two measurements are usually expressed as a ratio such as 100/80. These values are somewhat higher in older people. When blood pressures are elevated, such as 140/90, there is a greater risk of stroke, heart attack, or kidney damage. Low blood pressure prevents the brain from receiving adequate oxygen, causing dizziness and fainting.

The blood pressures are measured by a sphygmomanometer, an instrument consisting of a stethoscope and an inflatable cuff connected to a tube of mercury called a manometer. After the cuff is wrapped around the upper arm, it is pumped up with air until it cuts off the flow of blood through the arm. With the stethoscope over the artery, the air is slowly released from the cuff, decreasing the pressure on the artery. When the blood flow first starts again in the artery, a noise can be heard through the stethoscope signifying the systolic blood pressure as the pressure shown on the manometer. As air continues to be released, the cuff deflates until no sound is heard in the artery. A second pressure reading is taken at the moment of silence and denotes the diastolic pressure, the pressure when the heart is not contracting.

The use of digital blood pressure monitors is becoming more common. However, they have not been validated for use in all situations and can sometimes give inaccurate readings.

**Measurement of Gas Pressure**

When billions and billions of gas particles hit against the walls of a container, they exert pressure, which is a force acting on a certain area.

\[
\text{Pressure (} P \text{)} = \frac{\text{force}}{\text{area}}
\]

The atmospheric pressure can be measured using a barometer (see Figure 11.3). At a pressure of exactly 1 atmosphere (atm), a mercury column in an inverted tube would be exactly 760 mm high. One atmosphere (atm) is defined as exactly 760 mmHg (millimeters of mercury). One atmosphere is also 760 Torr, a pressure unit named to honor Evangelista Torricelli, the inventor of the barometer. Because the units of torr and mmHg are equal, they are used interchangeably. One atmosphere is also equivalent to 29.9 in. of mercury (inHg).

1 atm = 760 mmHg = 760 Torr (exact)
1 atm = 29.9 inHg
1 mmHg = 1 Torr (exact)

In SI units, pressure is measured in pascals (Pa); 1 atm is equal to 101,325 Pa. Because a pascal is a very small unit, pressures are usually reported in kilopascals.

1 atm = 101,325 Pa = 101.325 kPa

The U.S. equivalent of 1 atm is 14.7 lb/in.\(^2\) (psi). When you use a pressure gauge to check the air pressure in the tires of a car, it may read 30 to 35 psi. This measurement is actually 30 to 35 psi above the pressure that the atmosphere exerts on the outside of the tire.

1 atm = 14.7 lb/in.\(^2\)
**TABLE 11.2** summarizes the various units used in the measurement of pressure.

<table>
<thead>
<tr>
<th>Unit</th>
<th>Abbreviation</th>
<th>Unit Equivalent to 1 atm</th>
</tr>
</thead>
<tbody>
<tr>
<td>atmosphere</td>
<td>atm</td>
<td>1 atm (exact)</td>
</tr>
<tr>
<td>millimeters of Hg</td>
<td>mmHg</td>
<td>760 mmHg (exact)</td>
</tr>
<tr>
<td>torr</td>
<td>Torr</td>
<td>760 Torr (exact)</td>
</tr>
<tr>
<td>inches of Hg</td>
<td>inHg</td>
<td>29.9 inHg</td>
</tr>
<tr>
<td>pounds per square inch</td>
<td>lb/in.² (psi)</td>
<td>14.7 lb/in.²</td>
</tr>
<tr>
<td>pascal</td>
<td>Pa</td>
<td>101 325 Pa</td>
</tr>
<tr>
<td>kilopascal</td>
<td>kPa</td>
<td>101.325 kPa</td>
</tr>
</tbody>
</table>

Atmospheric pressure changes with variations in weather and altitude. On a hot, sunny day, the mercury column rises, indicating a higher atmospheric pressure. On a rainy day, the atmosphere exerts less pressure, which causes the mercury column to fall. In the weather report, this type of weather is called a low-pressure system. Above sea level, the density of the gases in the air decreases, which causes lower atmospheric pressures; the atmospheric pressure is greater than 760 mmHg at the Dead Sea because it is below sea level (see **TABLE 11.3**).

<table>
<thead>
<tr>
<th>Location</th>
<th>Altitude (km)</th>
<th>Atmospheric Pressure (mmHg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Dead Sea</td>
<td>−0.40</td>
<td>800</td>
</tr>
<tr>
<td>Sea level</td>
<td>0.00</td>
<td>760</td>
</tr>
<tr>
<td>Los Angeles</td>
<td>0.09</td>
<td>752</td>
</tr>
<tr>
<td>Las Vegas</td>
<td>0.70</td>
<td>700</td>
</tr>
<tr>
<td>Denver</td>
<td>1.60</td>
<td>630</td>
</tr>
<tr>
<td>Mount Whitney</td>
<td>4.50</td>
<td>440</td>
</tr>
<tr>
<td>Mount Everest</td>
<td>8.90</td>
<td>253</td>
</tr>
</tbody>
</table>

Divers must be concerned about increasing pressures on their ears and lungs when they dive below the surface of the ocean. Because water is more dense than air, the pressure on a diver increases rapidly as the diver descends. At a depth of 33 ft below the surface of the ocean, an additional 1 atm of pressure is exerted by the water on a diver, which gives a total pressure of 2 atm. At 100 ft, there is a total pressure of 4 atm on a diver. The regulator that a diver uses continuously adjusts the pressure of the breathing mixture to match the increase in pressure.

**SAMPLE PROBLEM 11.2** Units of Pressure

The oxygen in a tank in the hospital respiratory unit has a pressure of 4820 mmHg. Calculate the pressure, in atmospheres, of the oxygen gas.

**TRY IT FIRST**

**SOLUTION**

The equality $1 \text{ atm} = 760 \text{ mmHg}$ can be written as two conversion factors:

- $\frac{760 \text{ mmHg}}{1 \text{ atm}}$ and $\frac{1 \text{ atm}}{760 \text{ mmHg}}$
The anesthetic N₂O gas is used for pain relief.

**11.1 Properties of Gases**

**LEARNING GOAL** Describe the kinetic molecular theory of gases and the units of measurement used for gases.

11.1 Use the kinetic molecular theory of gases to explain each of the following:
- a. Gases move faster at higher temperatures.
- b. Gases can be compressed much more easily than liquids or solids.
- c. Gases have low densities.

11.2 Use the kinetic molecular theory of gases to explain each of the following:
- a. A container of nonstick cooking spray explodes when thrown into a fire.
- b. The air in a hot-air balloon is heated to make the balloon rise.
- c. You can smell the odor of cooking onions from far away.

11.3 Identify the property of a gas that is measured in each of the following:
- a. 350 K
- b. 125 mL
- c. 2.00 g of O₂
- d. 755 mmHg

11.4 Identify the property of a gas that is measured in each of the following:
- a. 425 K
- b. 1.0 atm
- c. 10.0 L
- d. 0.50 mol of He

**11.5** Which of the following statement(s) describes the pressure of a gas?
- a. the force of the gas particles on the walls of the container
- b. the number of gas particles in a container
- c. 4.5 L of helium gas
- d. 750 Torr
- e. 28.8 lb/in.²

**11.6** Which of the following statement(s) describes the pressure of a gas?
- a. 350 K
- b. the volume of the container
- c. 3.00 atm
- d. 0.25 mol of O₂
- e. 101 kPa

**Applications**

11.7 An tank contains oxygen (O₂) at a pressure of 2.00 atm. What is the pressure in the tank in terms of the following units?
- a. torr
- b. lb/in.²
- c. mmHg
- d. kPa

11.8 On a climb up Mount Whitney, the atmospheric pressure drops to 467 mmHg. What is the pressure in terms of the following units?
- a. atm
- b. torr
- c. inHg
- d. Pa
the volume \((V)\) of a sample of gas changes inversely with the pressure \((P)\) of the gas as long as there is no change in the temperature \((T)\) or amount of gas \((n)\), as illustrated in \textbf{FIGURE 11.4}.

If the volume or pressure of a gas changes without any change occurring in the temperature or in the amount of the gas, then the final pressure and volume will give the same \(PV\) product as the initial pressure and volume. Then we can set the initial and final \(PV\) products equal to each other. In the equation for Boyle’s law, the initial pressure and volume are written as \(P_1\) and \(V_1\) and the final pressure and volume are written as \(P_2\) and \(V_2\).

\begin{align*}
\text{Boyle’s Law} \quad & \quad P_1V_1 = P_2V_2 \\
\text{No change in temperature and number of moles} & \\
\end{align*}

\textbf{SAMPLE PROBLEM 11.3 Calculating Volume When Pressure Changes}

When Whitney had her asthma attack, she was given oxygen through a face mask. The gauge on a 12-L tank of compressed oxygen reads 3800 mmHg. How many liters would this same gas occupy at a final pressure of 570 mmHg when temperature and amount of gas do not change?

\textbf{TRY IT FIRST}

\textbf{SOLUTION}

\textbf{STEP 1} State the given and needed quantities. We place the gas data in a table by writing the initial pressure and volume as \(P_1\) and \(V_1\) and the final pressure and volume as \(P_2\) and \(V_2\). We see that the pressure decreases from 3800 mmHg to 570 mmHg. Using Boyle’s law, we predict that the volume increases.

\begin{table}
<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>(P_1 = 3800 \text{ mmHg})</td>
<td>(V_1 = 12 \text{ L})</td>
<td>Boyle’s law, (P_1V_1 = P_2V_2)</td>
</tr>
<tr>
<td>(P_2 = 570 \text{ mmHg})</td>
<td>(V_2)</td>
<td>Predict: (P) decreases, (V) increases</td>
</tr>
<tr>
<td>(\text{Factors that remain constant: } T \text{ and } n) &amp;</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

\textbf{STEP 2} Rearrange the gas law equation to solve for the unknown quantity. For a \(PV\) relationship, we use Boyle’s law and solve for \(V_2\) by dividing both sides by \(P_2\). According to Boyle’s law, a decrease in the pressure will cause an increase in the volume when \(T\) and \(n\) remain constant.

\begin{align*}
\frac{P_1V_1}{P_2} & = \frac{P_2V_2}{P_2} \\
\frac{P_1V_1}{P_2} & = \frac{P_2V_2}{P_2} \\
V_2 & = V_1 \times \frac{P_1}{P_2}
\end{align*}

\textbf{STEP 3} Substitute values into the gas law equation and calculate. When we substitute in the values with pressures in units of mmHg, the ratio of pressures (pressure factor) is greater than 1, which increases the volume as predicted.

\begin{align*}
V_2 & = 12 \text{ L} \times \frac{3800 \text{ mmHg}}{570 \text{ mmHg}} \\
& = 80. \text{ L}
\end{align*}

Pressure factor increases volume
STUDY CHECK 11.3
In an underground gas reserve, a bubble of methane gas (CH₄) has a volume of 45.0 mL at 1.60 atm pressure. What volume, in milliliters, will the gas bubble occupy when it reaches the surface where the atmospheric pressure is 744 mmHg, if there is no change in the temperature and amount of gas?

ANSWER
73.5 mL

CHEMISTRY LINK TO HEALTH
Pressure–Volume Relationship in Breathing

The importance of Boyle’s law becomes apparent when you consider the mechanics of breathing. Our lungs are elastic, balloon-like structures contained within an airtight chamber called the thoracic cavity. The diaphragm, a muscle, forms the flexible floor of the cavity.

Inspiration
The process of taking a breath of air begins when the diaphragm contracts and the rib cage expands, causing an increase in the volume of the thoracic cavity. The elasticity of the lungs allows them to expand when the thoracic cavity expands. According to Boyle’s law, the pressure inside the lungs decreases when their volume increases, causing the pressure inside the lungs to fall below the pressure of the atmosphere. This difference in pressures produces a pressure gradient between the lungs and the atmosphere. In a pressure gradient, molecules flow from an area of higher pressure to an area of lower pressure. During the inhalation phase of breathing, air flows into the lungs (inspiration), until the pressure within the lungs becomes equal to the pressure of the atmosphere.

Expiration
Expiration, or the exhalation phase of breathing, occurs when the diaphragm relaxes and moves back up into the thoracic cavity to its resting position. The volume of the thoracic cavity decreases, which squeezes the lungs and decreases their volume. Now the pressure in the lungs is higher than the pressure of the atmosphere, so air flows out of the lungs. Thus, breathing is a process in which pressure gradients are continuously created between the lungs and the environment because of the changes in the volume and pressure.

QUESTIONS AND PROBLEMS
11.2 Pressure and Volume (Boyle’s Law)

LEARNING GOAL Use the pressure–volume relationship (Boyle’s law) to calculate the unknown pressure or volume when the temperature and amount of gas are constant.

11.9 Why do scuba divers need to exhale air when they ascend to the surface of the water?

11.10 Why does a sealed bag of chips expand when you take it to a higher altitude?

11.11 The air in a cylinder with a piston has a volume of 220 mL and a pressure of 650 mmHg.

a. To obtain a higher pressure inside the cylinder at constant temperature and amount of gas, would the cylinder change as shown in A or B? Explain your choice.

b. If the pressure inside the cylinder increases to 1.2 atm, what is the final volume, in milliliters, of the cylinder?
11.12 A balloon is filled with helium gas. When each of the following changes are made at constant temperature, which of these diagrams (A, B, or C) shows the final volume of the balloon?

![Balloon Diagrams]

a. The balloon floats to a higher altitude where the outside pressure is lower.
b. The balloon is taken inside the house, but the atmospheric pressure does not change.
c. The balloon is put in a hyperbaric chamber in which the pressure is increased.

11.13 A gas with a volume of 4.0 L is in a closed container. Indicate the changes (increases, decreases, does not change) in its pressure when the volume undergoes the following changes at constant temperature and amount of gas:

a. The volume is compressed to 2.0 L.
b. The volume expands to 12 L.
c. The volume is compressed to 0.40 L.

11.14 A gas at a pressure of 2.0 atm is in a closed container. Indicate the changes (increases, decreases, does not change) in its volume when the pressure undergoes the following changes at constant temperature and amount of gas:

a. The pressure increases to 6.0 atm.
b. The pressure remains at 2.0 atm.
c. The pressure drops to 0.40 atm.

11.15 A 10.0-L balloon contains helium gas at a pressure of 655 mmHg. What is the final pressure, in millimeters of mercury, when the helium is placed in tanks that have the following volumes, if there is no change in temperature and amount of gas?

a. 20.0 L  
   b. 2.50 L  
   c. 13 800 mL  
   d. 1250 mL

11.16 The air in a 5.00-L tank has a pressure of 1.20 atm. What is the final pressure, in atmospheres, when the air is placed in tanks that have the following volumes, if there is no change in temperature and amount of gas?

a. 1.00 L  
   b. 2500 mL  
   c. 750 mL  
   d. 8.00 L

11.17 A sample of nitrogen (N₂) has a volume of 50.0 L at a pressure of 760. mmHg. What is the final volume, in liters, of the gas at each of the following pressures, if there is no change in temperature and amount of gas?

a. 725 mmHg  
   b. 2.0 atm  
   c. 0.500 atm  
   d. 850 Torr

11.18 A sample of methane (CH₄) has a volume of 25 mL at a pressure of 0.80 atm. What is the final volume, in milliliters, of the gas at each of the following pressures, if there is no change in temperature and amount of gas?

a. 0.40 atm  
   b. 2.00 atm  
   c. 2500 mmHg  
   d. 80.0 Torr

11.19 A sample of Ar gas has a volume of 5.40 L with an unknown pressure. The gas has a volume of 9.73 L when the pressure is 3.62 atm, with no change in temperature or amount of gas. What was the initial pressure, in atmospheres, of the gas?

11.20 A sample of Ne gas has a pressure of 654 mmHg with an unknown volume. The gas has a pressure of 345 mmHg when the volume is 495 mL, with no change in temperature or amount of gas. What was the initial volume, in milliliters, of the gas?

Applications

11.21 Cyclopropane, (C₃H₆), is a general anesthetic. A 5.0-L sample has a pressure of 5.0 atm. What is the final volume, in liters, of this gas given to a patient at a pressure of 1.0 atm with no change in temperature and amount of gas?

11.22 A tank holds 20.0 L of oxygen (O₂) at a pressure of 15.0 atm. What is the final volume, in liters, of this gas when it is released at a pressure of 1.00 atm with no change in temperature and amount of gas?

11.23 Use the words inspiration and expiration to describe the part of the breathing cycle that occurs as a result of each of the following:

a. The diaphragm contracts.
b. The volume of the lungs decreases.
c. The pressure within the lungs is less than that of the atmosphere.

11.24 Use the words inspiration and expiration to describe the part of the breathing cycle that occurs as a result of each of the following:

a. The diaphragm relaxes, moving up into the thoracic cavity.
b. The volume of the lungs expands.
c. The pressure within the lungs is higher than that of the atmosphere.

11.3 Temperature and Volume (Charles’s Law)

**LEARNING GOAL** Use the temperature–volume relationship (Charles’s law) to calculate the unknown temperature or volume when the pressure and amount of gas are constant.

Suppose that you are going to take a ride in a hot-air balloon. The captain turns on a propane burner to heat the air inside the balloon. As the air is heated, it expands and becomes less dense than the air outside, causing the balloon and its passengers to lift off. In 1787, Jacques Charles, a balloonist as well as a physicist, proposed that the volume of a gas is
related to the temperature. This proposal became Charles’s law, which states that the volume \( V \) of a gas is directly related to the temperature \( T \) when there is no change in the pressure \( P \) or amount \( n \) of gas. A direct relationship is one in which the related properties increase or decrease together. For two conditions, initial and final, we can write Charles’s law as follows:

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2}
\]

No change in pressure and number of moles

All temperatures used in gas law calculations must be converted to their corresponding Kelvin (K) temperatures.

To determine the effect of changing temperature on the volume of a gas, the pressure and the amount of gas are kept constant. If we increase the temperature of a gas sample, we know from the kinetic molecular theory that the motion (kinetic energy) of the gas particles will also increase. To keep the pressure constant, the volume of the container must increase (see Figure 11.5). If the temperature of the gas is decreased, the volume of the container must also decrease to maintain the same pressure when the amount of gas is constant.

**SAMPLE PROBLEM 11.4 Calculating Volume When Temperature Changes**

Helium gas is used to inflate the abdomen during laparoscopic surgery. A sample of helium gas has a volume of 5.40 L and a temperature of 15 °C. What is the final volume, in liters, of the gas after the temperature has been increased to 42 °C at constant pressure and amount of gas?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities. We place the gas data in a table by writing the initial temperature and volume as \( T_1 \) and \( V_1 \) and the final temperature and volume as \( T_2 \) and \( V_2 \). We see that the temperature increases from 15 °C to 42 °C. Using Charles’s law, we predict that the volume increases.

\[
\begin{align*}
T_1 &= 15 \, ^\circ\text{C} + 273 = 288 \, \text{K} \\
T_2 &= 42 \, ^\circ\text{C} + 273 = 315 \, \text{K}
\end{align*}
\]

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>( T_1 = 288 , \text{K} )</td>
<td>( T_2 = 315 , \text{K} )</td>
<td>Charles’s law, ( \frac{V_1}{T_1} = \frac{V_2}{T_2} )</td>
</tr>
<tr>
<td>( V_1 = 5.40 , \text{L} )</td>
<td>( V_2 )</td>
<td>Predict: ( T ) increases, ( V ) increases</td>
</tr>
<tr>
<td>Factors that remain constant: ( P ) and ( n )</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Rearrange the gas law equation to solve for the unknown quantity. In this problem, we want to know the final volume \( V_2 \) when the temperature increases. Using Charles’s law, we solve for \( V_2 \) by multiplying both sides by \( T_2 \).

\[
\begin{align*}
\frac{V_1}{T_1} &= \frac{V_2}{T_2} \\
V_2 &= V_1 \times \frac{T_2}{T_1}
\end{align*}
\]

Why does the volume of a gas increase when the temperature increases at constant pressure and amount of gas?
STEP 3 Substitute values into the gas law equation and calculate. From the table, we see that the temperature has increased. Because temperature is directly related to volume, the volume must increase. When we substitute in the values, we see that the ratio of the temperatures (temperature factor) is greater than 1, which increases the volume as predicted.

\[
\frac{V_2}{V_1} = \frac{315 \text{ K}}{288 \text{ K}} = 1.09
\]

Temperature factor increases volume

STUDY CHECK 11.4

A mountain climber inhales air that has a temperature of \(-8^\circ\text{C}\). If the final volume of air in the lungs is 569 mL at a body temperature of \(37^\circ\text{C}\), what was the initial volume of air, in milliliters, inhaled by the climber?

ANSWER

486 mL

QUESTIONS AND PROBLEMS

11.3 Temperature and Volume (Charles’s Law)

LEARNING GOAL Use the temperature–volume relationship (Charles’s law) to calculate the unknown temperature or volume when the pressure and amount of gas are constant.

11.25 Select the diagram that shows the final volume of a balloon when each of the following changes are made at constant pressure and amount of gas:

- a. The temperature is changed from 100 K to 300 K.
- b. The balloon is placed in a freezer.
- c. The balloon is first warmed and then returned to its starting temperature.

![Diagram of balloons and temperatures]

11.26 Indicate whether the final volume of gas in each of the following is the same, larger, or smaller than the initial volume, if pressure and amount of gas do not change:

- a. A volume of 505 mL of air on a cold winter day at \(-15^\circ\text{C}\) is breathed into the lungs, where body temperature is \(37^\circ\text{C}\).
- b. The heater used to heat the air in a hot-air balloon is turned off.
- c. A balloon filled with helium at the amusement park is left in a car on a hot day.

11.27 A sample of neon initially has a volume of 2.50 L at \(15^\circ\text{C}\). What final temperature, in degrees Celsius, is needed to change the volume of the gas to each of the following, if \(P\) and \(n\) do not change?

- a. 5.00 L
- b. 1250 mL
- c. 7.50 L
- d. 3550 mL

11.28 A gas has a volume of 4.00 L at \(0^\circ\text{C}\). What final temperature, in degrees Celsius, is needed to change the volume of the gas to each of the following, if \(P\) and \(n\) do not change?

- a. 1.50 L
- b. 1200 mL
- c. 10.0 L
- d. 50.0 mL

11.29 A balloon contains 2500 mL of helium gas at \(75^\circ\text{C}\). What is the final volume, in milliliters, of the gas when the temperature changes to each of the following, if \(P\) and \(n\) do not change?

- a. \(55^\circ\text{C}\)
- b. 680. K
- c. \(-25^\circ\text{C}\)
- d. 240. K

11.30 An air bubble has a volume of 0.500 L at \(18^\circ\text{C}\). What is the final volume, in liters, of the gas when the temperature changes to each of the following, if \(P\) and \(n\) do not change?

- a. \(0^\circ\text{C}\)
- b. 425 K
- c. \(-12^\circ\text{C}\)
- d. 575 K

11.31 A gas sample has a volume of 0.256 L with an unknown temperature. The same gas has a volume of 0.198 L when the temperature is \(32^\circ\text{C}\), with no change in the pressure or amount of gas. What was the initial temperature, in degrees Celsius, of the gas?

11.32 A gas sample has a temperature of \(22^\circ\text{C}\) with an unknown volume. The same gas has a volume of 456 mL when the temperature is \(86^\circ\text{C}\), with no change in the pressure or amount of gas. What was the initial volume, in milliliters, of the gas?
11.4 Temperature and Pressure (Gay-Lussac’s Law)

**LEARNING GOAL** Use the temperature–pressure relationship (Gay-Lussac’s law) to calculate the unknown temperature or pressure when the volume and amount of gas are constant.

If we could observe the molecules of a gas as the temperature rises, we would notice that they move faster and hit the sides of the container more often and with greater force. If we maintain a constant volume and amount of gas, the pressure would increase. In the temperature–pressure relationship known as Gay-Lussac’s law, the pressure of a gas is directly related to its Kelvin temperature. This means that an increase in temperature increases the pressure of a gas, and a decrease in temperature decreases the pressure of the gas as long as the volume and amount of gas do not change (see **Figure 11.6**).

**Gay-Lussac’s Law**

\[ \frac{P_1}{T_1} = \frac{P_2}{T_2} \quad \text{No change in volume and number of moles} \]

All temperatures used in gas law calculations must be converted to their corresponding Kelvin (K) temperatures.

**SAMPLE PROBLEM 11.5 Calculating Pressure When Temperature Changes**

Home oxygen tanks, which provide an oxygen-rich environment, can be dangerous if they are heated, because they can explode. Suppose an oxygen tank has a pressure of 120 atm at a room temperature of 25 °C. If a fire in the room causes the temperature of the gas inside the oxygen tank to reach 402 °C, what will be its pressure in atmospheres if the volume and amount of gas do not change? The oxygen tank may rupture if the pressure inside exceeds 180 atm. Would you expect it to rupture?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities. We place the gas data in a table by writing the initial temperature and pressure as \( T_1 \) and \( P_1 \) and the final temperature and pressure as \( T_2 \) and \( P_2 \). We see that the temperature increases from 25 °C to 402 °C. Using Gay-Lussac’s law, we predict that the pressure increases.

\[
T_1 = 25 ^\circ C + 273 = 298 K \\
T_2 = 402 ^\circ C + 273 = 675 K
\]

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>( P_1 = 4.0 \text{ atm} )</td>
<td>( P_2 )</td>
<td>Gay-Lussac’s law, ( \frac{P_1}{T_1} = \frac{P_2}{T_2} )</td>
</tr>
<tr>
<td>( T_1 = 298 \text{ K} )</td>
<td>( T_2 = 675 \text{ K} )</td>
<td>Predict: ( T ) increases, ( P ) increases</td>
</tr>
<tr>
<td>Factors that remain constant: ( V ) and ( n )</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Rearrange the gas law equation to solve for the unknown quantity. Using Gay-Lussac’s law, we solve for \( P_2 \) by multiplying both sides by \( T_2 \).

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2} \\
\frac{P_1}{T_1} \times T_2 = \frac{P_2}{T_2} \times T_2 \\
P_2 = P_1 \times \frac{T_2}{T_1}
\]
STEP 3 Substitute values into the gas law equation and calculate. When we substitute in the values, we see that the ratio of the temperatures (temperature factor) is greater than 1, which increases the pressure as predicted.

\[
P_2 = 120 \text{ atm} \times \frac{675 \text{ K}}{298 \text{ K}} = 270 \text{ atm}
\]

Temperature factor increases pressure

Because the calculated pressure of 270 atm exceeds the limit of 180 atm, we would expect the oxygen tank to rupture.

STUDY CHECK 11.5

In a storage area of a hospital where the temperature has reached 55 °C, the pressure of oxygen gas in a 15.0-L steel cylinder is 965 Torr. To what temperature, in degrees Celsius, would the gas have to be cooled to reduce the pressure to 850 Torr, when the volume and the amount of the gas do not change?

ANSWER

16 °C

Vapor Pressure and Boiling Point

When liquid molecules with sufficient kinetic energy break away from the surface, they become gas particles or vapor. In an open container, all the liquid will eventually evaporate. In a closed container, the vapor accumulates and creates pressure called vapor pressure. Each liquid exerts its own vapor pressure at a given temperature. As temperature increases, more vapor forms, and vapor pressure increases.

A liquid reaches its boiling point when its vapor pressure becomes equal to the external pressure. As boiling occurs, bubbles of the gas form within the liquid and quickly rise to the surface. For example, at an atmospheric pressure of 760 mmHg, water will boil at 100 °C, the temperature at which its vapor pressure reaches 760 mmHg (see Table 11.4).

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>Vapor Pressure (mmHg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>5</td>
</tr>
<tr>
<td>10</td>
<td>9</td>
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<td>20</td>
<td>18</td>
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<td>30</td>
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<td>37*</td>
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</tr>
<tr>
<td>100</td>
<td>760</td>
</tr>
</tbody>
</table>

*At body temperature
At high altitudes, where atmospheric pressures are lower than 760 mmHg, the boiling point of water is lower than 100 °C. For example, a typical atmospheric pressure in Denver is 630 mmHg. This means that water in Denver boils when the vapor pressure is 630 mmHg. Table 11.5 shows that an increase in the pressure for water increases the boiling point.

In a closed container such as a pressure cooker, a pressure greater than 1 atm can be obtained, which means that water boils at a temperature higher than 100 °C. Laboratories and hospitals use closed containers called autoclaves to sterilize laboratory and surgical equipment at temperature of 121 °C to 135 °C.

### Table 11.5 Pressure and the Boiling Point of Water

<table>
<thead>
<tr>
<th>Pressure (mmHg)</th>
<th>Boiling Point (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>270</td>
<td>70</td>
</tr>
<tr>
<td>467</td>
<td>87</td>
</tr>
<tr>
<td>630</td>
<td>95</td>
</tr>
<tr>
<td>752</td>
<td>99</td>
</tr>
<tr>
<td>760</td>
<td>100</td>
</tr>
<tr>
<td>800</td>
<td>100.4</td>
</tr>
<tr>
<td>1075</td>
<td>110</td>
</tr>
<tr>
<td>1520 (2 atm)</td>
<td>120</td>
</tr>
<tr>
<td>3800 (5 atm)</td>
<td>160</td>
</tr>
<tr>
<td>7600 (10 atm)</td>
<td>180</td>
</tr>
</tbody>
</table>

An autoclave used to sterilize equipment attains a temperature higher than 100 °C.

### QUESTIONS AND PROBLEMS

**11.4 Temperature and Pressure (Gay-Lussac’s Law)**

**Learning Goal** Use the temperature–pressure relationship (Gay-Lussac’s law) to calculate the unknown temperature or pressure when the volume and amount of gas are constant.

**11.33** Calculate the final pressure, in millimeters of mercury, for each of the following, if V and n do not change:

- **a.** A gas with an initial pressure of 1200 Torr at 155 °C is cooled to 0 °C.
- **b.** A gas in an aerosol can at an initial pressure of 1.40 atm at 12 °C is heated to 35 °C.

**11.34** Calculate the final pressure, in atmospheres, for each of the following, if V and n do not change:

- **a.** A gas with an initial pressure of 1.20 atm at 75 °C is cooled to −32 °C.
- **b.** A sample of N₂ with an initial pressure of 780. mmHg at −75 °C is heated to 28 °C.

**11.35** Calculate the final temperature, in degrees Celsius, for each of the following, if V and n do not change:

- **a.** A sample of xenon gas at 25 °C and 740. mmHg is cooled to give a pressure of 620. mmHg.
- **b.** A tank of argon gas with a pressure of 0.950 atm at −18 °C is heated to give a pressure of 1250 Torr.

**11.36** Calculate the final temperature, in degrees Celsius, for each of the following, if V and n do not change:

- **a.** A sample of helium gas with a pressure of 250 Torr at 0 °C is heated to give a pressure of 1500 Torr.
- **b.** A sample of air at 40 °C and 740. mmHg is cooled to give a pressure of 680. mmHg.

**11.37** A gas sample has a pressure of 766 mmHg with an unknown temperature. The same gas has a pressure of 744 mmHg when the temperature is 22 °C, with no change in the volume or amount of gas. What was the initial temperature, in degrees Celsius, of the gas?

**11.38** A gas sample has a unknown pressure with a temperature of 46 °C. The same gas has a pressure of 2.35 atm when the temperature is −15 °C, with no change in the volume or amount of gas. What was the initial pressure, in atmospheres, of the gas?

**11.39** Explain each of the following observations:

- **a.** Water boils at 87 °C on the top of Mount Whitney.
- **b.** Food cooks more quickly in a pressure cooker than in an open pan.

**11.40** Explain each of the following observations:

- **a.** Boiling water at sea level is hotter than boiling water in the mountains.
- **b.** Water used to sterilize surgical equipment is heated to 120 °C at 2.0 atm in an autoclave.

**Applications**

**11.41** A tank contains isoflurane, an inhaled anesthetic, at a pressure of 1.8 atm and 5 °C. What is the pressure, in atmospheres, if the gas is warmed to a temperature of 22 °C, if V and n do not change?

**11.42** Bacteria and viruses are inactivated by temperatures above 135 °C. An autoclave contains steam at 1.00 atm and 100 °C. At what pressure, in atmospheres, will the temperature of the steam in the autoclave reach 135 °C, if V and n do not change?
11.5 The Combined Gas Law

**LEARNING GOAL.** Use the combined gas law to calculate the unknown pressure, volume, or temperature of a gas when changes in two of these properties are given and the amount of gas is constant.

All of the pressure–volume–temperature relationships for gases that we have studied may be combined into a single relationship called the **combined gas law**. This expression is useful for studying the effect of changes in two of these variables on the third as long as the amount of gas (number of moles) remains constant.

**Combined Gas Law**

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
\]

No change in number of moles of gas

By using the combined gas law, we can derive any of the gas laws by omitting those properties that do not change, as seen in **Table 11.6**.

**Table 11.6 Summary of Gas Laws**

<table>
<thead>
<tr>
<th>Combined Gas Law</th>
<th>Properties Held Constant</th>
<th>Relationship</th>
<th>Name of Gas Law</th>
</tr>
</thead>
<tbody>
<tr>
<td>(\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2})</td>
<td>(T, n)</td>
<td>(P_1 V_1 = P_2 V_2)</td>
<td>Boyle’s law</td>
</tr>
<tr>
<td>(\frac{P V_1}{T_1} = \frac{P V_2}{T_2})</td>
<td>(P, n)</td>
<td>(V_1 = V_2)</td>
<td>Charles’s law</td>
</tr>
<tr>
<td>(\frac{P V_1}{T_1} = \frac{P V_2}{T_2})</td>
<td>(V, n)</td>
<td>(\frac{P_1}{T_1} = \frac{P_2}{T_2})</td>
<td>Gay-Lussac’s law</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 11.6 Using the Combined Gas Law**

A 25.0-mL bubble is released from a diver’s air tank at a pressure of 4.00 atm and a temperature of 11 °C. What is the volume, in milliliters, of the bubble when it reaches the ocean surface where the pressure is 1.00 atm and the temperature is 18 °C? (Assume the amount of gas in the bubble does not change.)

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities. We list the properties that change, which are the pressure, volume, and temperature. The property that remains constant, which is the amount of gas, is shown below the table. The temperatures in degrees Celsius must be changed to kelvins.

\[
T_1 = 11 ^\circ\text{C} + 273 = 284 \text{ K}
\]

\[
T_2 = 18 ^\circ\text{C} + 273 = 291 \text{ K}
\]

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>(P_1 = 4.00 \text{ atm})</td>
<td>(P_2 = 1.00 \text{ atm})</td>
<td>combined gas law, (\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2})</td>
</tr>
<tr>
<td>(V_1 = 25.0 \text{ mL})</td>
<td>(V_2)</td>
<td></td>
</tr>
<tr>
<td>(T_1 = 284 \text{ K})</td>
<td>(T_2 = 291 \text{ K})</td>
<td></td>
</tr>
<tr>
<td>Factor that remains constant: (n)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Under water, the pressure on a diver is greater than the atmospheric pressure.
**STEP 2** Rearrange the gas law equation to solve for the unknown quantity.

Using the combined gas law, we solve for $V_2$ by multiplying both sides by $T_2$ and dividing both sides by $P_2$.

$$
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
$$

Multiplying both sides by $T_2$ and dividing both sides by $P_2$:

$$
V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1}
$$

**STEP 3** Substitute values into the gas law equation and calculate. From the data table, we determine that both the pressure decrease and the temperature increase will increase the volume.

$$
V_2 = 25.0 \text{ mL} \times \frac{4.00 \text{ atm}}{1.00 \text{ atm}} \times \frac{291 \text{ K}}{284 \text{ K}} = 102 \text{ mL}
$$

However, in situations where the unknown value is decreased by one change but increased by the second change, it is difficult to predict the overall change for the unknown.

**STUDY CHECK 11.6**

A weather balloon is filled with 15.0 L of helium at a temperature of 25 °C and a pressure of 685 mmHg. What is the pressure, in millimeters of mercury, of the helium in the balloon in the upper atmosphere when the final temperature is −35 °C and the final volume becomes 34.0 L, if the amount of He does not change?

**ANSWER**

241 mmHg

**QUESTIONS AND PROBLEMS**

**11.5 The Combined Gas Law**

**LEARNING GOAL** Use the combined gas law to calculate the unknown pressure, volume, or temperature of a gas when changes in two of these properties are given and the amount of gas is constant.

11.43 Rearrange the variables in the combined gas law to solve for $T_2$.

11.44 Rearrange the variables in the combined gas law to solve for $P_2$.

11.45 A sample of helium gas has a volume of 6.50 L at a pressure of 845 mmHg and a temperature of 25 °C. What is the final pressure of the gas, in atmospheres, when the volume and temperature of the gas sample are changed to the following, if the amount of gas does not change?

- a. 1850 mL and 325 K
- b. 2.25 L and 12 °C
- c. 12.8 L and 47 °C

11.46 A sample of argon gas has a volume of 735 mL at a pressure of 1.20 atm and a temperature of 112 °C. What is the final volume of the gas, in milliliters, when the pressure and temperature of the gas sample are changed to the following, if the amount of gas does not change?

- a. 658 mmHg and 281 K
- b. 0.55 atm and 75 °C
- c. 15.4 atm and −15 °C

**Applications**

11.47 A 124-mL bubble of hot gas initially at 212 °C and 1.80 atm is emitted from an active volcano. What is the final temperature, in degrees Celsius, of the gas in the bubble outside the volcano if the final volume of the bubble is 138 mL and the pressure is 0.800 atm, if the amount of gas does not change?

11.48 A scuba diver 60 ft below the ocean surface inhales 50.0 mL of compressed air from a scuba tank at a pressure of 3.00 atm and a temperature of 8 °C. What is the final pressure of air, in atmospheres, in the lungs when the gas expands to 150.0 mL at a body temperature of 37 °C, if the amount of gas does not change?
11.6 Volume and Moles (Avogadro’s Law)

**LEARNING GOAL.** Use Avogadro’s law to calculate the unknown amount or volume of a gas when the pressure and temperature are constant.

In our study of the gas laws, we have looked at changes in properties for a specified amount \((n)\) of gas. Now we will consider how the properties of a gas change when there is a change in the number of moles or grams of the gas.

When you blow up a balloon, its volume increases because you add more air molecules. If the balloon has a small hole in it, air leaks out, causing its volume to decrease. In 1811, Amedeo Avogadro formulated **Avogadro’s law**, which states that the volume of a gas is directly related to the number of moles of a gas when temperature and pressure do not change. For example, if the number of moles of a gas is doubled, then the volume will also double as long as we do not change the pressure or the temperature (see **Figure 11.7**). At constant pressure and temperature, we can write Avogadro’s law as follows:

**Avogadro’s Law**

\[
\frac{V_1}{n_1} = \frac{V_2}{n_2}
\]

*No change in pressure and temperature*

**SAMPLE PROBLEM 11.7 Calculating Volume for a Change in Moles**

A weather balloon with a volume of 44 L is filled with 2.0 mol of helium. What is the final volume, in liters, if 3.0 mol of helium are added, to give a total of 5.0 mol of helium, if the pressure and temperature do not change?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities. We list those properties that change, which are volume and amount (moles). The properties that do not change, which are pressure and temperature, are shown below the table. Because there is an increase in the number of moles of gas, we can predict that the volume increases.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>(V_1 = 44) L (n_1 = 2.0) mol (n_2 = 5.0) mol</td>
<td>(V_2)</td>
<td>Avogadro’s law, (\frac{V_1}{n_1} = \frac{V_2}{n_2})</td>
<td>Predict: (n) increases, (V) increases</td>
</tr>
<tr>
<td>Factors that remain constant: (P) and (T)</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Rearrange the gas law equation to solve for the unknown quantity. Using Avogadro’s law, we can solve for \(V_2\) by multiplying both sides of the equation by \(n_2\).

\[
\frac{V_1}{n_1} = \frac{V_2}{n_2}
\]

\[
V_1 \times n_2 = V_2 \times n_2
\]

\[
V_2 = V_1 \times \frac{n_2}{n_1}
\]
STEP 3 Substitute values into the gas law equation and calculate. When we substitute in the values, we see that the ratio of the moles (mole factor) is greater than 1, which increases the volume as predicted.

\[ V_2 = 44 \text{ L} \times \frac{5.0 \text{ mol}}{2.0 \text{ mol}} = 110 \text{ L} \]

STUDY CHECK 11.7
A sample containing 8.00 g of oxygen gas has a volume of 5.00 L. What is the volume, in liters, after 4.00 g of oxygen gas is added to the 8.00 g of oxygen in the balloon, if the temperature and pressure do not change?

ANSWER
7.50 L

STP and Molar Volume
Using Avogadro’s law, we can say that any two gases will have equal volumes if they contain the same number of moles of gas at the same temperature and pressure. To help us make comparisons between different gases, arbitrary conditions called standard temperature (273 K) and standard pressure (1 atm) together abbreviated STP, were selected by scientists:

**STP Conditions**
- Standard temperature is exactly 0 °C (273 K).
- Standard pressure is exactly 1 atm (760 mmHg).

At STP, one mole of any gas occupies a volume of 22.4 L, which is about the same as the volume of three basketballs. This volume, 22.4 L, of any gas is called the **molar volume** (see Figure 11.8).

When a gas is at STP conditions (0 °C and 1 atm), its molar volume can be used to write conversion factors between the number of moles of gas and its volume, in liters.

**Molar Volume Conversion Factors**
- \( \frac{1 \text{ mol of gas}}{22.4 \text{ L (STP)}} = \frac{22.4 \text{ L (STP)}}{1 \text{ mol gas}} \)
SAMPLE PROBLEM 11.8 Using Molar Volume to Find Volume at STP

What is the volume, in liters, of 64.0 g of $O_2$ gas at STP?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>64.0 g of $O_2(g)$ at STP</td>
<td>liters of $O_2$ gas at STP</td>
<td>molar mass, molar volume (STP)</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to calculate the needed quantity.

grams of $O_2$ | Molar mass | moles of $O_2$ | Molar volume | liters of $O_2$

STEP 3 Write the equalities and conversion factors including 22.4 L/mol at STP.

$$
\frac{1 \text{ mol of } O_2 = 32.00 \text{ g of } O_2}{32.00 \text{ g } O_2} \quad \text{and} \quad \frac{1 \text{ mol } O_2}{1 \text{ mol } O_2} = \frac{22.4 \text{ L } O_2 \text{ (STP)}}{1 \text{ mol } O_2} \quad \text{and} \quad \frac{1 \text{ mol } O_2}{22.4 \text{ L } O_2 \text{ (STP)}}
$$

STEP 4 Set up the problem with factors to cancel units.

$$
64.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{22.4 \text{ L } O_2 \text{ (STP)}}{1 \text{ mol } O_2} = 44.8 \text{ L of } O_2 \text{ (STP)}
$$

STUDY CHECK 11.8

How many grams of $Cl_2(g)$ are in 5.00 L of $Cl_2(g)$ at STP?

ANSWER

15.8 g of $Cl_2(g)$

QUESTIONS AND PROBLEMS

11.6 Volume and Moles (Avogadro’s Law)

LEARNING GOAL Use Avogadro’s law to calculate the unknown amount or volume of a gas when the pressure and temperature are constant.

11.49 What happens to the volume of a bicycle tire or a basketball when you use an air pump to add air?

11.50 Sometimes when you blow up a balloon and release it, it flies around the room. What is happening to the air in the balloon and its volume?

11.51 A sample containing 1.50 mol of Ne gas has an initial volume of 8.00 L. What is the final volume, in liters, when each of the following changes occur in the quantity of the gas at constant pressure and temperature?

a. A leak allows one-half of Ne atoms to escape.
b. A sample of 3.50 mol of Ne is added to the 1.50 mol of Ne gas in the container.
c. A sample of 25.0 g of Ne is added to the 1.50 mol of Ne gas in the container.

11.52 A sample containing 4.80 g of $O_2$ gas has an initial volume of 15.0 L. What is the final volume, in liters, when each of the following changes occur in the quantity of the gas at constant pressure and temperature?

a. A sample of 0.500 mol of $O_2$ is added to the 4.80 g of $O_2$ in the container.
b. A sample of 2.00 g of $O_2$ is removed.
c. A sample of 4.00 g of $O_2$ is added to the 4.80 g of $O_2$ gas in the container.

11.53 Use the molar volume to calculate each of the following at STP:

a. the number of moles of $O_2$ in 44.8 L of $O_2$ gas
b. the volume, in liters, occupied by 2.50 mol of $N_2$ gas
c. the volume, in liters, occupied by 50.0 g of Ar gas
d. the number of grams of $H_2$ in 1620 mL of $H_2$ gas

11.54 Use the molar volume to calculate each of the following at STP:

a. the number of moles of $CO_2$ in 4.00 L of $CO_2$ gas
b. the volume, in liters, occupied by 0.420 mol of He gas
c. the volume, in liters, occupied by 6.40 g of $O_2$ gas
d. the number of grams of Ne contained in 11.2 L of Ne gas
11.7 The Ideal Gas Law

**LEARNING GOAL** Use the ideal gas law equation to calculate the unknown \( P, V, T, \) or \( n \) of a gas when given three of the four values in the ideal gas law equation. Calculate the molar mass of a gas.

The ideal gas law is the relationship between the four properties used in the measurement of a gas—pressure \( (P) \), volume \( (V) \), temperature \( (T) \), and amount of a gas \( (n) \).

**Ideal Gas Law**

\[ PV = nRT \]

Rearranging the ideal gas law equation shows that the four gas properties equal the gas law constant, \( R \).

\[ \frac{PV}{nT} = R \]

To calculate the value of \( R \), we substitute the STP conditions for molar volume into the expression: 1.00 mol of any gas occupies 22.4 L at STP (273 K and 1.00 atm).

\[ R = \frac{(1.00 \text{ atm})(22.4 \text{ L})}{(1.00 \text{ mol})(273 \text{ K})} = \frac{0.0821 \text{ L \cdot atm}}{\text{mol \cdot K}} \]

The value for the ideal gas constant, \( R \), is 0.0821 L \cdot atm per mol \cdot K. If we use 760. mmHg for the pressure, we obtain another useful value for \( R \) of 62.4 L \cdot mmHg per mol \cdot K.

\[ R = \frac{(760. \text{ mmHg})(22.4 \text{ L})}{(1.00 \text{ mol})(273 \text{ K})} = \frac{62.4 \text{ L \cdot mmHg}}{\text{mol \cdot K}} \]

The ideal gas law is a useful expression when you are given the quantities for any three of the four properties of a gas. Although real gases show some deviations in behavior, the ideal gas law closely approximates the behavior of real gases at typical conditions. In working problems using the ideal gas law, the units of each variable must match the units in the \( R \) you select.

<table>
<thead>
<tr>
<th>Ideal Gas Component</th>
<th>Units</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ideal Gas Constant ((R))</td>
<td>0.0821 L \cdot atm \</td>
</tr>
<tr>
<td>Pressure ((P))</td>
<td>atm</td>
</tr>
<tr>
<td>Volume ((V))</td>
<td>L</td>
</tr>
<tr>
<td>Amount ((n))</td>
<td>mol</td>
</tr>
<tr>
<td>Temperature ((T))</td>
<td>K</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 11.9 Using the Ideal Gas Law**

Dinitrogen oxide, \( \text{N}_2\text{O} \), which is an anesthetic also called laughing gas, is used in dentistry. What is the pressure, in atmospheres, of 0.350 mol of \( \text{N}_2\text{O} \) at 22 °C in a 5.00-L container?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities. When three of the four quantities \( (P, V, n, \) and \( T) \) are known, we use the ideal gas law equation to solve for the unknown quantity. It is helpful to organize the data in a table. The temperature is converted from degrees Celsius to kelvins so that the units of \( V, n, \) and \( T \) match the unit of the gas constant \( R \).

Dinitrogen oxide is used as an anesthetic in dentistry.
Guide to Using the Ideal Gas Law

**STEP 1** State the given and needed quantities.

**STEP 2** Rearrange the ideal gas law equation to solve for the needed quantity.

**STEP 3** Substitute the gas data into the equation and calculate the needed quantity.

---

**SAMPLE PROBLEM 11.10 Calculating Mass Using the Ideal Gas Law**

Butane, C₄H₁₀, is used as a fuel for camping stoves. If you have 108 mL of butane gas at 715 mmHg and 25 °C, what is the mass, in grams, of butane?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities. When three of the quantities (P, V, and T) are known, we use the ideal gas law equation to solve for the quantity moles (n). The volume given in milliliters (mL) is converted to a volume in liters (L).

---

**STUDY CHECK 11.9**

Chlorine gas, Cl₂, is used to purify water. How many moles of chlorine gas are in a 7.00-L tank if the gas has a pressure of 865 mmHg and a temperature of 24 °C?

**ANSWER**

0.327 mol of Cl₂

Many times we need to know the amount of gas, in grams. Then the ideal gas law equation can be rearranged to solve for the amount (n) of gas, which is converted to mass in grams using its molar mass as shown in Sample Problem 11.10.

---

When camping, butane is used as a fuel for a portable stove.
**STEP 2** Rearrange the ideal gas law equation to solve for the needed quantity.

By dividing both sides of the ideal gas law equation by \( RT \), we solve for moles, \( n \).

\[
PV = nRT \quad \text{(Ideal gas law equation)}
\]

\[
\frac{PV}{RT} = \frac{nRT}{RT}
\]

\[
n = \frac{PV}{RT}
\]

**STEP 3** Substitute the gas data into the equation and calculate the needed quantity.

\[
n = \frac{715 \text{ mmHg} \times 0.108 \text{ L}}{62.4 \text{ L·mmHg·mol}^{-1} \cdot \text{K} \times 298 \text{ K}} = 0.00415 \text{ mol} \times (4.15 \times 10^{-3} \text{ mol})
\]

Now we can convert the moles of butane to grams using its molar mass of 58.12 g/mol:

\[
0.00415 \text{ mol C}_4\text{H}_{10} \times \frac{58.12 \text{ g C}_4\text{H}_{10}}{1 \text{ mol C}_4\text{H}_{10}} = 0.241 \text{ g of C}_4\text{H}_{10}
\]

**STUDY CHECK 11.10**

What is the volume, in liters, of 1.20 g of carbon monoxide at 8 °C if it has a pressure of 724 mmHg?

**ANSWER**

1.04 L

**Molar Mass of a Gas**

Another use of the ideal gas law is to determine the molar mass of a gas. If the mass, in grams, of the gas is known, we calculate the number of moles of the gas using the ideal gas law equation. Then the molar mass (g/mol) can be determined.

**SAMPLE PROBLEM 11.11 Molar Mass of a Gas Using the Ideal Gas Law**

What is the molar mass, in grams per mole, of a gas if a 3.16-g sample of gas at 0.750 atm and 45 °C occupies a volume of 2.05 L?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>( P = 0.750 \text{ atm} )</td>
<td>( n )</td>
<td>molar mass, ( PV = nRT )</td>
</tr>
<tr>
<td>( V = 2.05 \text{ L} )</td>
<td></td>
<td>( R = \frac{0.0821 \text{ L·atm}}{\text{mol·K}} )</td>
</tr>
<tr>
<td>( T = 45 \text{ °C} + 273 = 318 \text{ K} )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>mass = 3.16 g</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Guide to Calculating the Molar Mass of a Gas**

**STEP 1** State the given and needed quantities.

**STEP 2** Rearrange the ideal gas law equation to solve for the number of moles.

**STEP 3** Obtain the molar mass by dividing the given number of grams by the number of moles.
**STEP 2** Rearrange the ideal gas law equation to solve for the number of moles.

To solve for moles, \( n \), divide both sides of the ideal gas law equation by \( RT \).

\[
PV = nRT \quad \text{Ideal gas law equation}
\]

\[
\frac{PV}{RT} = \frac{nRT}{RT}
\]

\[
\frac{PV}{RT} = n
\]

\[
 n = \frac{PV}{RT}
\]

**STEP 3** Obtain the molar mass by dividing the given number of grams by the number of moles.

\[
\text{Molar mass} = \frac{\text{mass}}{\text{moles}}
\]

**STUDY CHECK 11.11**

What is the molar mass, in grams per mole, of an unknown gas in a 1.50-L container if 0.488 g of the gas has a pressure of 0.0750 atm at 19 °C?

**ANSWER**

104 g/mol

### QUESTIONS AND PROBLEMS

#### 11.7 The Ideal Gas Law

**LEARNING GOAL** Use the ideal gas law equation to calculate the unknown \( P, V, T, \) or \( n \) of a gas when given three of the four values in the ideal gas law equation. Calculate the molar mass of a gas.

11.55 Calculate the pressure, in atmospheres, of 2.00 mol of helium gas in a 10.0-L container at 27 °C.

11.56 What is the volume, in liters, of 4.00 mol of methane gas, \( \text{CH}_4 \), at 18 °C and 1.40 atm?

11.57 An oxygen gas container has a volume of 20.0 L. How many grams of oxygen are in the container if the gas has a pressure of 845 mmHg at 22 °C?

11.58 A 10.0-g sample of krypton has a temperature of 25 °C at 575 mmHg. What is the volume, in milliliters, of the krypton gas?

11.59 A 25.0-g sample of nitrogen, \( \text{N}_2 \), has a volume of 50.0 L and a pressure of 630. mmHg. What is the temperature, in kelvins, and degrees Celsius, of the gas?

11.60 A 0.226-g sample of carbon dioxide, \( \text{CO}_2 \), has a volume of 525 mL and a pressure of 455 mmHg. What is the temperature, in kelvins and degrees Celsius, of the gas?

11.61 Determine the molar mass of each of the following gases:

a. 0.84 g of a gas that occupies 450 mL at 0 °C and 1.00 atm (STP)

b. 1.28 g of a gas that occupies 1.00 L at 0 °C and 760 mmHg (STP)

c. 1.48 g of a gas that occupies 1.00 L at 685 mmHg and 22 °C

d. 2.96 g of a gas that occupies 2.30 L at 0.95 atm and 24 °C

11.62 Determine the molar mass of each of the following gases:

a. 2.90 g of a gas that occupies 0.500 L at 0 °C and 1.00 atm (STP)

b. 1.43 g of a gas that occupies 2.00 L at 0 °C and 760 mmHg (STP)

c. 0.726 g of a gas that occupies 855 mL at 1.20 atm and 18 °C

d. 2.32 g of a gas that occupies 1.23 L at 685 mmHg and 25 °C

#### Applications

11.63 A single-patient hyperbaric chamber has a volume of 640 L. At a temperature of 24 °C, how many grams of oxygen are needed to give a pressure of 1.6 atm?

11.64 A multipatient hyperbaric chamber has a volume of 3400 L. At a temperature of 22 °C, how many grams of oxygen are needed to give a pressure of 2.4 atm?
11.8 Gas Laws and Chemical Reactions

LEARNING GOAL  Calculate the mass or volume of a gas that reacts or forms in a chemical reaction.

Gases are involved as reactants and products in many chemical reactions. For example, we have seen that the combustion of organic fuels with oxygen gas produces carbon dioxide gas and water vapor. In combination reactions, we have seen that hydrogen gas and nitrogen gas react to form ammonia gas, and hydrogen gas and oxygen gas produce water. Typically, the information given for a gas in a reaction is its pressure ($P$), volume ($V$), and temperature ($T$). Then we can use the ideal gas law equation to determine the moles of a gas in a reaction. If we are given the number of moles for one of the gases in a reaction, we can use a mole–mole factor to determine the moles of any other substance.

SAMPLE PROBLEM 11.12 Gases in Chemical Reactions

Limestone (CaCO$_3$) reacts with HCl to produce carbon dioxide gas, water, and aqueous calcium chloride.

$$\text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{CaCl}_2(aq)$$

How many liters of CO$_2$ are produced at 752 mmHg and 24 °C from a 25.0-g sample of limestone?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>25.0 g of CaCO$_3$</td>
<td>$V$ of CO$_2$(g)</td>
<td>ideal gas law, $PV = nRT$</td>
</tr>
<tr>
<td>$P = 752$ mmHg</td>
<td>$T = 24 , ^\circ$ C $+ 273 = 297$ K</td>
<td>$R = \frac{62.4 , L \cdot \text{mmHg}}{\text{mol} \cdot \text{K}}$</td>
</tr>
<tr>
<td></td>
<td>$R = \frac{62.4 , L \cdot \text{mmHg}}{\text{mol} \cdot \text{K}}$</td>
<td>molar mass</td>
</tr>
</tbody>
</table>

Equation

$$\text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{CaCl}_2(aq)$$

STEP 2 Write a plan to convert the given quantity to the needed moles.

grams of CaCO$_3$ Molar mass moles of CaCO$_3$ Mole–mole factor $\rightarrow$ moles of CO$_2$

STEP 3 Write the equalities and conversion factors for molar mass and mole–mole factors.

$$\frac{1 \, \text{mol of CaCO}_3}{100.09 \, \text{g CaCO}_3} \quad \text{and} \quad \frac{1 \, \text{mol CaCO}_3}{100.09 \, \text{g CaCO}_3}$$

$$\frac{1 \, \text{mol of CaCO}_3}{1 \, \text{mol of CO}_2} \quad \text{and} \quad \frac{1 \, \text{mol CO}_2}{1 \, \text{mol CaCO}_3}$$

STEP 4 Set up the problem to calculate moles of needed quantity.

$$25.0 \, \text{g CaCO}_3 \times \frac{1 \, \text{mol CaCO}_3}{100.09 \, \text{g CaCO}_3} \times \frac{1 \, \text{mol CO}_2}{1 \, \text{mol CaCO}_3} = 0.250 \, \text{mol of CO}_2$$
**11.8 Gas Laws and Chemical Reactions**

**LEARNING GOAL** Calculate the mass or volume of a gas that reacts or forms in a chemical reaction.

11.65 Mg metal reacts with HCl to produce hydrogen gas.

\[ \text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{H}_2(g) + \text{MgCl}_2(aq) \]

a. What volume, in liters, of hydrogen at 0 °C and 1.00 atm (STP) is released when 8.25 g of Mg reacts?

b. How many grams of magnesium are needed to prepare 5.00 L of H₂ at 735 mmHg and 19 °C?

11.66 When heated to 350 °C at 0.950 atm, ammonium nitrate decomposes to produce nitrogen, water, and oxygen gases.

\[ 2\text{NH}_4\text{NO}_3(s) \rightarrow 2\text{N}_2(g) + 4\text{H}_2\text{O}(g) + \text{O}_2(g) \]

a. How many liters of water vapor are produced when 25.8 g of NH₄NO₃ decomposes?

b. How many grams of NH₄NO₃ are needed to produce 10.0 L of oxygen?

11.67 Butane undergoes combustion when it reacts with oxygen to produce carbon dioxide and water. What volume, in liters, of oxygen is needed to react with 55.2 g of butane at 0.850 atm and 25 °C?

\[ 2\text{C}_4\text{H}_{10}(g) + 13\text{O}_2(g) \rightarrow 8\text{CO}_2(g) + 10\text{H}_2\text{O}(g) \]

11.68 Potassium nitrate decomposes to potassium nitrite and oxygen. What volume, in liters, of O₂ can be produced from the decomposition of 50.0 g of KNO₃ at 35 °C and 1.19 atm?

\[ 2\text{KNO}_3(s) \rightarrow 2\text{KNO}_2(s) + \text{O}_2(g) \]

11.69 Aluminum and oxygen react to form aluminum oxide. How many liters of oxygen at 0 °C and 760 mmHg (STP) are required to completely react with 5.4 g of aluminum?

\[ 4\text{Al}(s) + 3\text{O}_2(g) \rightarrow 2\text{Al}_2\text{O}_3(s) \]

11.70 Nitrogen dioxide reacts with water to produce oxygen and ammonia. How many grams of NH₃ can be produced when 4.00 L of NO₂ reacts at 415 °C and 725 mmHg?

\[ 4\text{NO}_2(g) + 6\text{H}_2\text{O}(g) \rightarrow 7\text{O}_2(g) + 4\text{NH}_3(g) \]

**11.9 Partial Pressures (Dalton’s Law)**

**LEARNING GOAL** Use Dalton’s law of partial pressures to calculate the total pressure of a mixture of gases.

Many gas samples are a mixture of gases. For example, the air you breathe is a mixture of mostly oxygen and nitrogen gases. In ideal gas mixtures, scientists observed that all gas particles behave in the same way. Therefore, the total pressure of the gases in a mixture is a result of the collisions of the gas particles regardless of what type of gas they are.

In a gas mixture, each gas exerts its partial pressure, which is the pressure it would exert if it were the only gas in the container. Dalton’s law states that the total pressure of a gas mixture is the sum of the partial pressures of the gases in the mixture.
Dalton’s Law

\[ P_{\text{total}} = P_1 + P_2 + P_3 + \cdots \]

Total pressure of a gas mixture = Sum of the partial pressures of the gases in the mixture

Suppose we have two separate tanks, one filled with helium at a pressure of 2.0 atm and the other filled with argon at a pressure of 4.0 atm. When the gases are combined in a single tank with the same volume and temperature, the number of gas molecules, not the type of gas, determines the pressure in the container. There the pressure of the gas mixture would be 6.0 atm, which is the sum of their individual or partial pressures.

![Image of two gas tanks with pressures 2.0 atm and 4.0 atm, combined to show total pressure of 6.0 atm.]

The total pressure of two gases is the sum of their partial pressures.

Air Is a Gas Mixture

The air you breathe is a mixture of gases. What we call the atmospheric pressure is actually the sum of the partial pressures of the gases in the air. Table 11.7 lists partial pressures for the gases in air on a typical day.

### Table 11.7 Typical Composition of Air

<table>
<thead>
<tr>
<th>Gas</th>
<th>Partial Pressure (mmHg)</th>
<th>Percentage (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrogen, N(_2)</td>
<td>594</td>
<td>78.2</td>
</tr>
<tr>
<td>Oxygen, O(_2)</td>
<td>160</td>
<td>21.0</td>
</tr>
<tr>
<td>Carbon dioxide, CO(_2)</td>
<td>6</td>
<td>0.8</td>
</tr>
<tr>
<td>Argon, Ar</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Water vapor, H(_2)O</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Total air</td>
<td>760</td>
<td>100</td>
</tr>
</tbody>
</table>

### Sample Problem 11.13 Partial Pressure of a Gas in a Mixture

A heliox breathing mixture of oxygen and helium is prepared for a patient with chronic obstructive pulmonary disease (COPD). The gas mixture has a total pressure of 7.00 atm. If the partial pressure of the oxygen in the tank is 1140 mmHg, what is the partial pressure, in atmospheres, of the helium in the breathing mixture?

**Solution**

**Step 1** Write the equation for the sum of the partial pressures.

\[ P_{\text{total}} = P_O + P_{\text{He}} \]

**Step 2** Rearrange the equation to solve for the unknown pressure. To solve for the partial pressure of helium \(P_{\text{He}}\), we rearrange the equation to give the following:

\[ P_{\text{He}} = P_{\text{total}} - P_O \]
Convert units to match.

\[ P_{O_2} = \frac{1140 \text{ mmHg}}{760 \text{ mmHg}} \times 1 \text{ atm} = 1.50 \text{ atm} \]

**STEP 3** Substitute known pressures into the equation and calculate the unknown pressure.

\[ P_{He} = P_{\text{total}} - P_{O_2} \]
\[ P_{He} = 7.00 \text{ atm} - 1.50 \text{ atm} = 5.50 \text{ atm} \]

**STUDY CHECK 11.13**

An anesthetic consists of a mixture of cyclopropane gas, \( C_3H_6 \), and oxygen gas, \( O_2 \). If the mixture has a total pressure of 1.09 atm, and the partial pressure of the cyclopropane is 73 mmHg, what is the partial pressure, in millimeters of mercury, of the oxygen in the anesthetic?

**ANSWER**

755 mmHg

**Gases Collected Over Water**

In the laboratory, gases are often collected by bubbling them through water into a container (see **FIGURE 11.9**). In a reaction, magnesium (Mg) reacts with HCl to form H\(_2\) gas and MgCl\(_2\).

\[
\text{Mg}(s) + 2\text{HCl}(aq) \rightarrow \text{H}_2(g) + \text{MgCl}_2(aq)
\]

As hydrogen is produced during the reaction, it displaces some of the water in the container. Because of the vapor pressure of water, the gas that is collected is a mixture of hydrogen and water vapor. For our calculation, we need the pressure of the dry hydrogen gas. We use the vapor pressure of water (see Table 11.4) at the experimental temperature, and subtract it from the total gas pressure. Then we can use the ideal gas law to determine the moles or grams of the hydrogen gas that were collected.

![FIGURE 11.9](image) A gas from a reaction is collected by bubbling through water. Due to evaporation of water, the total pressure is equal to the partial pressure of the gas and the vapor pressure of water.

**How is the pressure of the dry gas determined?**
SAMPLE PROBLEM 11.14 Moles of Gas Collected Over Water

When magnesium reacts with HCl, a volume of 0.355 L of hydrogen gas is collected over water at 26 °C. The vapor pressure of water at 26 °C is 25 mmHg.

\[ \text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{H}_2(g) + \text{MgCl}_2(aq) \]

If the total pressure is 752 mmHg, how many moles of \( \text{H}_2(g) \) were collected?

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>Analysis</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>( V = 0.355 \text{ L of } \text{H}_2 )</td>
<td>moles of ( \text{H}_2 )</td>
<td>ideal gas law, ( PV = nRT )</td>
<td></td>
</tr>
<tr>
<td>( P = 752 \text{ mmHg} )</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>( T = 26 ^\circ \text{C} + 273 = 299 \text{ K} )</td>
<td></td>
<td>( R = \frac{62.4 \text{ L} \cdot \text{mmHg}}{\text{mol} \cdot \text{K}} )</td>
<td></td>
</tr>
<tr>
<td>( P_{\text{H}_2\text{O}} = 25 \text{ mmHg} )</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Obtain the vapor pressure of water. The vapor pressure of water at 26 °C is 25 mmHg.

**STEP 2** Subtract the vapor pressure from the total gas pressure to give the partial pressure of the needed gas. Using Dalton’s law of partial pressures, determine the partial pressure of \( \text{H}_2 \).

\[ P_{\text{total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}} \]

Solving for the partial pressure of \( \text{H}_2 \) gives

\[ P_{\text{H}_2} = P_{\text{total}} - P_{\text{H}_2\text{O}} \]

\[ P_{\text{H}_2} = 752 \text{ mmHg} - 25 \text{ mmHg} \]

\[ P_{\text{H}_2} = 727 \text{ mmHg} \]

**STEP 3** Use the ideal gas law to convert \( P_{\text{gas}} \) to moles of gas collected. By dividing both sides of the ideal gas law equation by \( RT \), we solve for moles, \( n \), of gas.

\[ PV = nRT \]

\[ \frac{PV}{RT} = \frac{nRT}{RT} \]

\[ n = \frac{PV}{RT} \]

Calculate the moles of \( \text{H}_2 \) gas by placing the partial pressure of \( \text{H}_2 \) (727 mmHg), volume of gas container (0.355 L), temperature (26 °C + 273 = 299 K), and \( R \), using mmHg, into the ideal gas law equation.

\[ n = \frac{727 \text{ mmHg} \times 0.355 \text{ L}}{62.4 \text{ L} \cdot \text{mmHg} \cdot \text{mol} \cdot \text{K}} \times 299 \text{ K} \]

\[ n = 0.0138 \text{ mol of } \text{H}_2 \]

**STUDY CHECK 11.14**

A 456-mL sample of oxygen gas (\( \text{O}_2 \)) is collected over water at a pressure of 744 mmHg and a temperature of 20 °C. How many grams of dry \( \text{O}_2 \) are collected?

**ANSWER**

0.579 g of \( \text{O}_2 \)
A burn patient may undergo treatment for burns and infections in a hyperbaric chamber, a device in which pressures can be obtained that are two to three times greater than atmospheric pressure. A greater oxygen pressure increases the level of dissolved oxygen in the blood and tissues, where it fights bacterial infections. High levels of oxygen are toxic to many strains of bacteria. The hyperbaric chamber may also be used during surgery, to help counteract carbon monoxide (CO) poisoning, and to treat some cancers.

The blood is normally capable of dissolving up to 95% of the oxygen. Thus, if the partial pressure of the oxygen in the hyperbaric chamber is 2280 mmHg (3 atm), about 2170 mmHg of oxygen can dissolve in the blood, saturating the tissues. In the treatment for carbon monoxide poisoning, oxygen at high pressure is used to displace the CO from the hemoglobin faster than breathing pure oxygen at 1 atm.

A patient undergoing treatment in a hyperbaric chamber must also undergo decompression (reduction of pressure) at a rate that slowly reduces the concentration of dissolved oxygen in the blood. If decompression is too rapid, the oxygen dissolved in the blood may form gas bubbles in the circulatory system.

Similarly, if a scuba diver does not decompress slowly, a condition called the “bends” may occur. While below the surface of the ocean, a diver uses a breathing mixture with higher pressures. If there is nitrogen in the mixture, higher quantities of nitrogen gas will dissolve in the blood. If the diver ascends to the surface too quickly, the dissolved nitrogen forms gas bubbles that can block a blood vessel and cut off the flow of blood in the joints and tissues of the body and be quite painful. A diver suffering from the bends is placed immediately into a hyperbaric chamber where pressure is first increased and then slowly decreased. The dissolved nitrogen can then diffuse through the lungs until atmospheric pressure is reached.

**CHEMISTRY LINK TO HEALTH**

**Hyperbaric Chambers**

A hyperbaric chamber is used in the treatment of certain diseases.

**QUESTIONS AND PROBLEMS**

**11.9 Partial Pressures (Dalton’s Law)**

**LEARNING GOAL** Use Dalton’s law of partial pressures to calculate the total pressure of a mixture of gases.

11.71 In a gas mixture, the partial pressures are nitrogen 425 Torr, oxygen 115 Torr, and helium 225 Torr. What is the total pressure, in torr, exerted by the gas mixture?

11.72 In a gas mixture, the partial pressures are argon 415 mmHg, neon 75 mmHg, and nitrogen 125 mmHg. What is the total pressure, in millimeters of mercury, exerted by the gas mixture?

11.73 An air sample in the lungs contains oxygen at 93 mmHg, nitrogen at 565 mmHg, carbon dioxide at 38 mmHg, and water vapor at 47 mmHg. What is the total pressure, in atmospheres, exerted by the gas mixture?

11.74 A nitrox II gas mixture for scuba diving contains oxygen gas at 53 atm and nitrogen gas at 94 atm. What is the total pressure, in torr, of the scuba gas mixture?

11.75 A gas mixture containing oxygen, nitrogen, and helium exerts a total pressure of 925 Torr. If the partial pressures are oxygen 425 Torr and helium 75 Torr, what is the partial pressure, in torr, of the nitrogen in the mixture?

11.76 A gas mixture containing oxygen, nitrogen, and neon exerts a total pressure of 1.20 atm. If helium added to the mixture increases the pressure to 1.50 atm, what is the partial pressure, in atmospheres, of the helium?

11.77 When solid KClO₃ is heated, it decomposes to give solid KCl and O₂ gas. A volume of 256 mL of gas is collected over water at a total pressure of 765 mmHg and 24 °C. The vapor pressure of water at 24 °C is 22 mmHg.

\[2\text{KClO}_3(s) \xrightarrow{\Delta} 2\text{KCl}(s) + 3\text{O}_2(g)\]

a. What was the partial pressure of the O₂ gas?

b. How many moles of O₂ gas were produced in the reaction?

11.78 When Zn reacts with HCl solution, the products are H₂ gas and ZnCl₂. A volume of 425 mL of H₂ gas is collected over water at a total pressure of 758 mmHg and 16 °C. The vapor pressure of water at 16 °C is 14 mmHg.

\[\text{Zn}(s) + \text{HCl}(aq) \longrightarrow \text{H}_2(g) + \text{ZnCl}_2(aq)\]

a. What was the partial pressure of the H₂ gas?

b. How many moles of H₂ gas were produced in the reaction?
Follow Up

**EXERCISE-INDUCED ASTHMA**

Vigorous exercise with high levels of physical activity can induce asthma particularly in children. When Whitney had her asthma attack, her breathing became more rapid, the temperature within her airways increased, and the muscles around the bronchi contracted, causing a narrowing of the airways. Whitney’s symptoms, which may occur within 5 to 20 min after the start of vigorous exercise, include shortness of breath, wheezing, and coughing.

Whitney now does several things to prevent exercise-induced asthma. She uses a pre-exercise inhaled medication before she starts her activity. The medication relaxes the muscles that surround the airways and opens up the airways. Then she does a warm-up set of exercises. If pollen counts are high, she avoids exercising outdoors.

**Applications**

11.79 Whitney’s lung capacity was measured as 3.2 L at a body temperature of 37 °C and a pressure of 745 mmHg. How many moles of oxygen are in her lungs if air contains 21% oxygen?

11.80 Whitney’s tidal volume, which is the volume of air that she inhales and exhales, was 0.54 L. Her tidal volume was measured at a body temperature of 37 °C and a pressure of 745 mmHg. How many moles of nitrogen does she inhale in one breath if air contains 78% nitrogen?

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**CONCEPT MAP**

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**GASES**

- Behave according to **Gas Laws**
  - **P and V** Boyle’s Law
  - **V and T** Charles’s Law
  - **P and T** Gay-Lussac’s Law
  - **P, V, and T** Combined Gas Law

- Gas mixture

- Volume and moles are related by **Avogadro’s Law**

- Molar Volume

- Partial Pressures

- Reactions of Gases

- Molar Mass of a Gas
11.1 Properties of Gases

LEARNING GOAL Describe the kinetic molecular theory of gases and the units of measurement used for gases.

- In a gas, particles are so far apart and moving so fast that their attractions are negligible.
- A gas is described by the physical properties of pressure \( (P) \), volume \( (V) \), temperature \( (T) \), and amount in moles \( (n) \).
- A gas exerts pressure, the force of the gas particles striking the surface of a container.
- Gas pressure is measured in units such as torr, mmHg, atm, and Pa.

11.2 Pressure and Volume (Boyle’s Law)

LEARNING GOAL Use the pressure–volume relationship (Boyle’s law) to calculate the unknown pressure or volume when the temperature and amount of gas are constant.

- The volume \( (V) \) of a gas changes inversely with the pressure \( (P) \) of the gas if there is no change in the temperature and the amount of gas.
  \[ P_1V_1 = P_2V_2 \]
- The pressure of a gas increases if its volume decreases; its pressure decreases if the volume increases.

11.3 Temperature and Volume (Charles’s Law)

LEARNING GOAL Use the temperature–volume relationship (Charles’s law) to calculate the unknown temperature or volume when the pressure and amount of gas are constant.

- The volume \( (V) \) of a gas is directly related to its Kelvin temperature \( (T) \) when there is no change in the pressure and the amount of gas.
  \[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]
- If the temperature of a gas increases, its volume increases; if its temperature decreases, the volume decreases.

11.4 Temperature and Pressure (Gay-Lussac’s Law)

LEARNING GOAL Use the temperature–pressure relationship (Gay-Lussac’s law) to calculate the unknown temperature or pressure when the volume and amount of gas are constant.

- The pressure \( (P) \) of a gas is directly related to its Kelvin temperature \( (T) \).
  \[ \frac{P_1}{T_1} = \frac{P_2}{T_2} \]
- An increase in temperature increases the pressure of a gas, and a decrease in temperature decreases the pressure, if there is no change in the volume and the amount of gas.
- Vapor pressure is the pressure of the gas that forms when a liquid evaporates.
- At the boiling point of a liquid, the vapor pressure equals the external pressure.

11.5 The Combined Gas Law

LEARNING GOAL Use the combined gas law to calculate the unknown pressure, volume, or temperature of a gas when changes in two of these properties are given and the amount of gas is constant.

- The combined gas law is the relationship of pressure \( (P) \), volume \( (V) \), and temperature \( (T) \) for a constant amount of gas.
  \[ \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \]
- The combined gas law is used to determine the effect of changes in two of the variables on the third.

11.6 Volume and Moles (Avogadro’s Law)

LEARNING GOAL Use Avogadro’s law to calculate the unknown amount or volume of a gas when the pressure and temperature are constant.

- The volume \( (V) \) of a gas is directly related to the number of moles \( (n) \) of the gas when the pressure and temperature of the gas do not change.
  \[ \frac{V_1}{n_1} = \frac{V_2}{n_2} \]
- If the moles of gas increase, the volume must increase; if the moles of gas decrease, the volume must decrease.
- At standard temperature (273 K) and standard pressure (1 atm), abbreviated STP, 1 mol of any gas has a volume of 22.4 L.

11.7 The Ideal Gas Law

LEARNING GOAL Use the ideal gas law equation to calculate the unknown \( P, V, T, \) or \( n \) of a gas when given three of the four values in the ideal gas law equation. Calculate the molar mass of a gas.

- The ideal gas law can be used to calculate the mass, number of moles, or the volume of a gas in a reaction.
- The ideal gas law gives the relationship of the quantities \( P, V, n, \) and \( T \) that describe and measure a gas.
  \[ PV = nRT \]
- Any of the four variables can be calculated if the values of the other three are known.
- The molar mass of a gas can be calculated using molar volume at STP or the ideal gas law.
### 11.8 Gas Laws and Chemical Reactions

**LEARNING GOAL** Calculate the mass or volume of a gas that reacts or forms in a chemical reaction.

- The ideal gas law equation is used to convert the quantities \( P, V, \) and \( T \) of gases to moles in a chemical reaction.
- The moles of gases can be used to determine the number of moles or grams of other substances in the reaction.

### 11.9 Partial Pressures (Dalton’s Law)

**LEARNING GOAL** Use Dalton’s law of partial pressures to calculate the total pressure of a mixture of gases.

- In a mixture of two or more gases, the total pressure is the sum of the partial pressures of the individual gases.

\[
P_{\text{total}} = P_1 + P_2 + P_3 + \ldots
\]

- The partial pressure of a gas in a mixture is the pressure it would exert if it were the only gas in the container.
- For gases collected over water, the vapor pressure of water is subtracted from the total pressure of the gas mixture to obtain the partial pressure of the dry gas.

### KEY TERMS

- **atmosphere (atm)** A unit equal to the pressure exerted by a column of mercury 760 mm high.
- **atmospheric pressure** The pressure exerted by the atmosphere.
- **Avogadro’s law** A gas law stating that the volume of a gas is directly related to the number of moles of gas when pressure and temperature do not change.
- **Boyle’s law** A gas law stating that the pressure of a gas is inversely related to the volume when temperature and moles of the gas do not change.
- **Charles’s law** A gas law stating that the volume of a gas is directly related to the Kelvin temperature when pressure and moles of the gas do not change.
- **combined gas law** A relationship that combines several gas laws relating pressure, volume, and temperature when the amount of gas does not change.

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
\]

- **Dalton’s law** A gas law stating that the total pressure exerted by a mixture of gases in a container is the sum of the partial pressures that each gas would exert alone.
- **direct relationship** A relationship in which two properties increase or decrease together.

### CORE CHEMISTRY SKILLS

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

#### Using the Gas Laws (11.2)

- Boyle’s, Charles’s, Gay-Lussac’s, and Avogadro’s laws show the relationships between two properties of a gas.

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \text{Boyle’s law}
\]

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{Charles’s law}
\]

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2} \quad \text{Gay-Lussac’s law}
\]

\[
\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{Avogadro’s law}
\]

- The combined gas law shows the relationship between \( P, V, \) and \( T \) for a gas.

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
\]

- When two properties of a gas vary and the other two are constant, we list the initial and final conditions of each property in a table.

**Example:** A sample of helium gas (He) has a volume of 6.8 L and a pressure of 2.5 atm. What is the final volume, in liters, if it has a final pressure of 1.2 atm with no change in temperature and amount of gas?
Answer:

Using Boyle’s law, we can write the relationship for \( V_2 \), which we predict will increase.

\[
V_2 = V_1 \times \frac{P_1}{P_2} = 6.8 \text{ L} \times \frac{2.5 \text{ atm}}{1.2 \text{ atm}} = 14 \text{ L}
\]

Using the Ideal Gas Law (11.7)

- The ideal gas law equation combines the relationships of the four properties of a gas into one equation.

\[
P V = n R T
\]

- When three of the four properties are given, we rearrange the ideal gas law equation for the needed quantity.

Example: What is the volume, in liters, of 0.750 mol of CO\(_2\) at a pressure of 1340 mmHg and a temperature of 295 K?

Answer: \[
V = \frac{n R T}{P} = \frac{0.750 \text{ mol} \times 62.4 \text{ L} \cdot \text{mmHg}^{-1} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 295 \text{ K}}{1340 \text{ mmHg}} = 10.3 \text{ L}
\]

Calculating Mass or Volume of a Gas in a Chemical Reaction (11.8)

- The ideal gas law equation is used to calculate the volume or mass of a gas in a chemical reaction.

Example: What is the volume, in liters, of N\(_2\) required to react with 18.5 g of magnesium at a pressure of 1.20 atm and a temperature of 303 K?

\[
3\text{Mg(s)} + \text{N}_2(g) \rightarrow \text{Mg}_3\text{N}_2(s)
\]

Answer: Initially, we convert the grams of Mg to moles and use a mole–mole factor from the balanced equation to calculate the moles of N\(_2\) gas.

\[
18.5 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{1 \text{ mol N}_2}{3 \text{ mol Mg}} = 0.254 \text{ mol of N}_2
\]

Now, we use the ideal gas law equation and solve for liters, the needed quantity.

\[
V = \frac{n R T}{P} = \frac{0.254 \text{ mol-N}_2 \times 0.0821 \text{ L} \cdot \text{atm}^{-1} \cdot \text{mol}^{-1} \times 303 \text{ K}}{1.20 \text{ atm}} = 5.27 \text{ L}
\]

Calculating Partial Pressure (11.9)

- In a gas mixture, each gas exerts its partial pressure, which is the pressure it would exert if it were the only gas in the container.

- Dalton’s law states that the total pressure of a gas mixture is the sum of the partial pressures of the gases in the mixture.

\[
P_{\text{total}} = P_1 + P_2 + P_3 + \cdots
\]

Example: A gas mixture with a total pressure of 1.18 atm contains helium gas at a partial pressure of 465 mmHg and nitrogen gas. What is the partial pressure, in atmospheres, of the nitrogen gas?

Answer: Initially, we convert the partial pressure of helium gas from mmHg to atm.

\[
465 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.612 \text{ atm of He gas}
\]

Using Dalton’s law, we solve for the needed quantity, \( P_{\text{N}_2} \).

\[
P_{\text{total}} = P_{\text{N}_2} + P_{\text{He}}
\]

\[
P_{\text{N}_2} = P_{\text{total}} - P_{\text{He}} = 1.18 \text{ atm} - 0.612 \text{ atm} = 0.57 \text{ atm}
\]

**UNDERSTANDING THE CONCEPTS**

*The chapter sections to review are shown in parentheses at the end of each question.*

**11.81** Two balloons of equal volume and at the same temperature contain different gases. One balloon contains 100 cm\(^3\) of He and the other balloon contains 100 cm\(^3\) of Ar. Is each of the following statements true or false? Explain. (11.1)

a. Both balloons are of the same weight.

b. Both balloons contain the same number of moles of particles.

**11.82** Two flasks of equal volume and at the same temperature contain different gases. One flask contains 10.0 g of N\(_2\), and the other flask contains 10.0 g of Cl\(_2\). Is each of the following statements true or false? Explain. (11.1)

a. The densities of the gases are the same.

b. The pressures in the flasks are different.

**11.83** With an increase in temperature from 20 °C to 80 °C, which of the following diagrams represents a gas sample with the (11.1)

a. smallest volume?  

b. largest volume?
11.84 Indicate which diagram (1, 2, or 3) represents the volume of the gas sample in a flexible container when each of the following changes (a to e) takes place: (11.2, 11.3)

![Diagrams](image)

- a. Temperature increases at constant pressure.
- b. Temperature decreases at constant pressure.
- c. Atmospheric pressure decreases at constant temperature.
- d. Doubling the atmospheric pressure and doubling the Kelvin temperature.

11.85 A balloon is filled with helium gas with a partial pressure of 1.00 atm and neon gas with a partial pressure of 0.50 atm. For each of the following changes (a to e) of the initial balloon, select the diagram (A, B, or C) that shows the final volume of the balloon: (11.2, 11.3, 11.6)

![Diagrams](image)

- a. The balloon is put in a cold storage unit \((P \text{ and } n \text{ constant})\).
- b. The balloon floats to a higher altitude where the pressure is less \((T \text{ and } n \text{ constant})\).
- c. All of the helium gas is removed \((T \text{ and } P \text{ constant})\).
- d. The Kelvin temperature doubles, and half of the gas atoms leak out \((P \text{ constant})\).
- e. 2.0 mol of \(O_2\) gas is added \((T \text{ and } P \text{ constant})\).

11.86 Indicate if pressure increases, decreases, or stays the same in each of the following: (11.2, 11.4, 11.6)

- a. 
- b. 
- c. 

11.87 An airplane is pressurized to 650 mmHg. (11.9)

- a. If air is 21% oxygen, what is the partial pressure of oxygen on the plane?
- b. If the partial pressure of oxygen drops below 100 mmHg, passengers become drowsy. If this happens, oxygen masks are released. What is the total cabin pressure at which oxygen masks are dropped?

11.88 At a restaurant, a customer chokes on a piece of food. You put your arms around the person’s waist and use your fists to push up on the person’s abdomen, an action called the Heimlich maneuver. (11.2)

- a. How would this action change the volume of the chest and lungs?
- b. Why does it cause the person to expel the food item from the airway?

11.89 In 1783, Jacques Charles launched his first balloon filled with hydrogen gas, which he chose because it was lighter than air. The balloon had a volume of 31 000 L when it was filled at a pressure of 755 mmHg and a temperature of 22 °C. When the balloon reached an altitude of 1000 m, the pressure was 658 mmHg and the temperature was −8 °C. What is the volume of the balloon at these conditions, if the amount of hydrogen remains the same? (11.5, 11.6)

11.90 Your spaceship has docked at a space station above Mars. The temperature inside the space station is a carefully controlled 24 °C at a pressure of 745 mmHg. A balloon with a volume of 425 mL drifts into the airlock where the temperature is −95 °C and the pressure is 0.115 atm. What is the final volume, in milliliters, of the balloon if \(n\) does not change and the balloon is very elastic? (11.5)

11.91 A sample of xenon gas has a volume of 7.50 L at 35 °C and 675 mmHg. What is the final pressure of the gas, in atmospheres, when the volume and temperature of the gas sample are changed to 18.5 L and 53 °C, if the amount of gas does not change? (11.5)

11.92 A weather balloon has a volume of 750 L when filled with helium at 8 °C at a pressure of 380 Torr. What is the final volume, in liters, of the balloon when the pressure is 0.20 atm, the temperature is −45 °C, and \(n\) does not change? (11.5)

11.93 A 10.00 L fuel tank is filled with acetylene gas, \(C_2H_2\), at a pressure of 1000 mmHg and a temperature of 30 °C. How many grams of acetylene are in the fuel tank? (11.7)
11.94 An iron container with a volume of 20.0 L is filled with 80.0 g of oxygen gas at 30 °C. What is the pressure, in atmospheres, of the O₂ gas in the cylinder? (11.7)

11.95 A sample of gas with a mass of 1.62 g occupies a volume of 941 mL at a pressure of 748 Torr and a temperature of 20.0 °C. What is the molar mass of the gas? (11.7)

11.96 What is the molar mass of a gas if 1.15 g of the gas has a volume of 8 mL at 0 °C and 1.00 atm (STP)? (11.7)

11.97 How many grams of SO₂ are in 40.0 L of SO₂(g) at 2.40 atm and 10 °C? (11.7)

11.98 A container is filled with 0.324 g of N₂ at 8 °C and 625 mmHg. What is the volume, in milliliters, of the container? (11.7)

11.99 How many liters of H₂ gas can be produced at 0 °C and 1.00 atm (STP) from 25.0 g of Zn? (11.7, 11.8)

11.100 In the formation of smog, nitrogen and oxygen gas react to form nitrogen dioxide. How many grams of NO₂ will be produced when 2.0 L of nitrogen at 840 mmHg and 24 °C are completely reacted? (11.7, 11.8)

11.101 Nitrogen dioxide reacts with water to produce oxygen and ammonia. A 5.00-L sample of H₂O(g) reacts at a temperature of 375 °C and a pressure of 725 mmHg. How many grams of NH₃ can be produced? (11.7, 11.8)

11.102 Hydrogen gas can be produced in the laboratory through the reaction of magnesium metal with hydrochloric acid. When 12.0 g of Mg reacts, what volume, in liters, of H₂ gas is produced at 24 °C and 835 mmHg? (11.7, 11.8)

11.103 A gas mixture with a total pressure of 2400 Torr is used by a scuba diver. If the mixture contains 2.0 mol of helium and 6.0 mol of oxygen, what is the partial pressure, in torr, of each gas in the sample? (11.9)

11.104 A gas mixture with a total pressure of 4.6 atm is used in a hospital. If the mixture contains 5.4 mol of nitrogen and 1.4 mol of oxygen, what is the partial pressure, in atmospheres, of each gas in the sample? (11.9)

11.105 A gas mixture contains oxygen and argon at partial pressures of 0.60 atm and 425 mmHg. If nitrogen gas added to the sample increases the total pressure to 2500 Torr, what is the partial pressure, in torr, of the nitrogen added? (11.9)

11.106 What is the total pressure, in millimeters of mercury, of a gas mixture containing argon gas at 0.25 atm, helium gas at 350 mmHg, and nitrogen gas at 360 Torr? (11.9)

---

### CHALLENGE QUESTIONS

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

11.107 Solid aluminum reacts with aqueous H₂SO₄ to form H₂ gas and aluminum sulfate. When a sample of Al is allowed to react, 415 mL of gas is collected over water at 23 °C, at a pressure of 755 mmHg. At 23 °C, the vapor pressure of water is 21 mmHg. (11.7, 11.8, 11.9)

   \[2Al(s) + 3H₂SO₄(aq) → 3H₂(g) + Al₂(SO₄)₃(aq)\]

   a. What is the pressure, in millimeters of mercury, of the dry H₂ gas?
   b. How many moles of H₂ were produced?
   c. How many grams of Al were reacted?

11.108 When heated, KClO₃ forms KCl and O₂. When a sample of KClO₃ is heated, 226 mL of gas with a pressure of 744 mmHg is collected over water at 26 °C. At 26 °C, the vapor pressure of water is 25 mmHg. (11.7, 11.8, 11.9)

   \[2KClO₃(s) → 2KCl(s) + 3O₂(g)\]

   a. What is the pressure, in millimeters of mercury, of the dry O₂ gas?
   b. How many moles of O₂ were produced?
   c. How many grams of KClO₃ were reacted?

11.109 A sample of gas with a mass of 1.020 g occupies a volume of 762 mL at 0 °C and 1.00 atm (STP). (7.4, 7.5, 11.6, 11.7)

   a. What is the molar mass of the gas?
   b. If the unknown gas is composed of 0.815 g of carbon and the rest is hydrogen, what is its molecular formula?

11.110 A sample of an unknown gas with a mass of 3.24 g occupies a volume of 1.88 L at a pressure of 748 mmHg and a temperature of 20 °C. (7.4, 7.5, 11.6, 11.7)

   a. What is the molar mass of the gas?
   b. If the unknown gas is composed of 2.78 g of carbon and the rest is hydrogen, what is its molecular formula?

11.111 Ethanol, C₂H₅OH, undergoes combustion with oxygen in the air to produce CO₂ and H₂O. How many liters of CO₂ are produced at STP on burning 610 g of ethanol? (11.7)

   \[C₂H₅OH(g) + 3O₂(g) → 2CO₂(g) + 3H₂O(g)\]

11.112 When sensors in a car detect a collision, they cause the reaction of sodium azide, NaN₃, which generates nitrogen gas to fill the airbags within 0.03 s. How many liters of N₂ are produced at STP if the airbag contains 132 g of NaN₃? (11.6, 11.7)

   \[2NaN₃(s) → 2Na(s) + 3N₂(g)\]

11.113 Glucose, C₆H₁₂O₆, is metabolized in living systems. How many grams of water can be produced from the reaction of 18.0 g of glucose and 7.50 L of O₂ at 1.00 atm and 37 °C? (9.2, 9.3, 11.7, 11.8)

   \[C₆H₁₂O₆(s) + 6O₂(g) → 6CO₂(g) + 6H₂O(l)\]

11.114 A 15.0 g piece of zinc metal is added to dilute hydrochloric acid and the hydrogen gas evolved is collected under atmospheric pressure at 20 °C. Assuming all of the zinc is reacted and zinc is the limiting reagent, what is the expected volume of hydrogen gas? (9.2, 9.3, 11.7, 11.8)

   \[Zn(s) + 2HCl(aq) → ZnCl₂(aq) + H₂(g)\]
Answers to Selected Questions and Problems

11.1 a. At a higher temperature, gas particles have greater kinetic energy, which makes them move faster.
   b. Because there are great distances between the particles of a gas, they can be pushed closer together and still remain a gas.
   c. Gas particles are very far apart, which means that the mass of a gas in a certain volume is very small, resulting in a low density.

11.3 a. temperature           b. volume
   c. amount                  d. pressure

11.5 Statements a, d, and e describe the pressure of a gas.

11.7 a. 1520 Torr
   b. 29.4 lb/in.²
   c. 1520 mmHg
   d. 203 kPa

11.9 As a diver ascends to the surface, external pressure decreases. If the air in the lungs were not exhaled, its volume would expand and severely damage the lungs. The pressure in the lungs must adjust to changes in the external pressure.

11.11 a. The pressure is greater in cylinder A. According to Boyle’s law, a decrease in volume pushes the gas particles closer together, which will cause an increase in the pressure.
   b. 160 mL

11.13 a. increases           b. decreases
   c. increases

11.15 a. 328 mmHg
   b. 2620 mmHg
   c. 475 mmHg
   d. 5240 mmHg

11.17 a. 52.4 L
   b. 25 L
   c. 100. L
   d. 45 L

11.19 6.52 atm

11.21 25 L of cyclopropane

11.23 a. inspiration
   b. expiration
   c. inspiration

11.25 a. C
   b. A
   c. B

11.27 a. 303 °C
   b. −129 °C
   c. 591 °C
   d. 136 °C

11.29 a. 2400 mL
   b. 4900 mL
   c. 1800 mL
   d. 1700 mL

11.31 121 °C

11.33 a. 770 mmHg
   b. 1150 mmHg

11.35 a. −23 °C
   b. 168 °C

11.37 31 °C

11.39 a. On top of a mountain, water boils below 100 °C because the atmospheric (external) pressure is less than 1 atm.
   b. Because the pressure inside a pressure cooker is greater than 1 atm, water boils above 100 °C. At a higher temperature, food cooks faster.

11.41 1.9 atm

11.43 \( T_2 = T_1 \times \frac{P_2}{P_1} \times \frac{V_1}{V_2} \)

11.45 a. 4.26 atm
   b. 3.07 atm
   c. 0.606 atm

11.47 −33 °C

11.49 The volume increases because the number of gas particles is increased.

11.51 a. 4.00 L
   b. 26.7 L
   c. 14.6 L

11.53 a. 2.00 mol of O₂
   b. 56.0 L
   c. 28.0 L
   d. 0.146 g of H₂

11.55 4.93 atm

11.57 29.4 g of O₂

11.59 566 K (293 °C)

11.61 a. 42 g/mol
   b. 28.7 g/mol
   c. 39.8 g/mol
   d. 33 g/mol

11.63 1300 g of O₂

11.65 a. 7.60 L of H₂
   b. 4.92 g of Mg

11.67 178 L of O₂

11.69 3.4 L of O₂

11.71 765 Torr

11.73 0.978 atm

11.75 425 Torr

11.77 a. 743 mmHg
   b. 0.0103 mol of O₂

11.79 0.025 mol of O₂

11.81 a. False. Both balloons contain different gases and each gas has its own molecular weight, thus having different weights.
   b. True. As the balloons are of the same volume, they have the same number of moles of particles.

11.83 a. 2
   b. 3

   b. C: Volume increases when pressure decreases.
   c. A: Volume decreases when the moles of gas decrease.
   d. B: Doubling the temperature, in kelvins, would double the volume, but when half of the gas escapes, the volume would decrease by half. These two opposing effects cancel each other, and there is no change in the volume.
   e. C: Increasing the moles increases the volume to keep T and P constant.

11.87 a. 140 mmHg
   b. 480 mmHg

11.89 32 000 L

11.91 0.383 atm

11.93 13.8 g of C₂H₂

11.95 42.1 g/mol

11.97 265 g

11.99 8.56 L of H₂

11.101 1.02 g of NH₃

11.103 He 600 Torr, O₂ 1800 Torr

11.105 370 Torr

11.107 a. 734 mmHg
   b. 0.0165 mol of H₂
   c. 0.297 g of Al

11.109 a. 30.0 g/mol
   b. C₂H₆

11.111 594 L of CO₂

11.113 5.31 g of water
**OUR KIDNEYS PRODUCE**

Urine, which carries waste products and excess fluid from the body. They also reabsorb electrolytes such as potassium and produce hormones that regulate blood pressure and the levels of calcium in the blood. Diseases such as diabetes and high blood pressure can cause a decrease in kidney function. Symptoms of kidney malfunction include protein in the urine, an abnormal level of urea nitrogen in the blood, frequent urination, and swollen feet. If kidney failure occurs, it may be treated with dialysis or transplantation.

Michelle has been suffering from kidney disease because of severe strep throat she contracted as a child. When her kidneys stopped functioning, Michelle was placed on dialysis three times a week. As she enters the dialysis unit, her dialysis nurse, Amanda, asks Michelle how she is feeling. Michelle indicates that she feels tired today and has considerable swelling around her ankles. Amanda informs her that these side effects occur because of her body’s inability to regulate the amount of water in her cells. Amanda explains that the amount of water is regulated by the concentration of electrolytes in her body fluids and the rate at which waste products are removed from her body. Amanda explains that although water is essential for the many chemical reactions that occur in the body, the amount of water can become too high or too low because of various diseases and conditions. Because Michelle’s kidneys no longer perform dialysis, she cannot regulate the amount of electrolytes or waste in her body fluids. As a result, she has an electrolyte imbalance and a buildup of waste products, so her body is retaining water. Amanda then explains that the dialysis machine does the work of her kidneys to reduce the high levels of electrolytes and waste products.

**CAREER**

**Dialysis Nurse**

A dialysis nurse specializes in assisting patients with kidney disease undergoing dialysis. This requires monitoring the patient before, during, and after dialysis for any complications such as a drop in blood pressure or cramping. The dialysis nurse connects the patient to the dialysis unit via a dialysis catheter that is inserted into the neck or chest, which must be kept clean to prevent infection. A dialysis nurse must have considerable knowledge about how the dialysis machine functions to ensure that it is operating correctly at all times.
Solutions are everywhere around us. Most of the gases, liquids, and solids we see are mixtures of at least one substance dissolved in another. There are different types of solutions. The air we breathe is a solution that is primarily oxygen and nitrogen gases. Carbon dioxide gas dissolved in water makes carbonated drinks. When we make solutions of coffee or tea, we use hot water to dissolve substances from coffee beans or tea leaves. The ocean is also a solution, consisting of many ionic compounds such as sodium chloride dissolved in water. In your medicine cabinet, the antiseptic tincture of iodine is a solution of iodine dissolved in ethanol. A solution is a homogeneous mixture in which one substance, called the solute, is uniformly dispersed in another substance called the solvent. Because the solute and the solvent do not react with each other, they can be mixed in varying proportions. A solution of a little salt dissolved in water tastes slightly salty. When a large amount of salt is dissolved in water, the solution tastes very salty. Usually, the solute (in this case, salt) is the substance present in the lesser amount, whereas the solvent (in this case, water) is present in the greater amount. For example, in a solution composed of 5.0 g of salt and 50.0 g of water, salt is the solute and water is the solvent. In a solution, the particles of the solute are evenly dispersed among the molecules within the solvent (see **FIGURE 12.1**).
Types of Solutes and Solvents

Solutes and solvents may be solids, liquids, or gases. The solution that forms has the same physical state as the solvent. When sugar crystals are dissolved in water, the resulting sugar solution is liquid. Sugar is the solute, and water is the solvent. Soda water and soft drinks are prepared by dissolving carbon dioxide gas in water. The carbon dioxide gas is the solute, and water is the solvent. **TABLE 12.1** lists some solutes and solvents and their solutions.

<table>
<thead>
<tr>
<th>Type</th>
<th>Example</th>
<th>Primary Solute</th>
<th>Solvent</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Gas Solutions</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Gas in a gas</td>
<td>Air</td>
<td>O₂(g)</td>
<td>N₂(g)</td>
</tr>
<tr>
<td><strong>Liquid Solutions</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Gas in a liquid</td>
<td>Soda water</td>
<td>CO₂(g)</td>
<td>H₂O(l)</td>
</tr>
<tr>
<td>Household ammonia</td>
<td>NH₃(g)</td>
<td>H₂O(l)</td>
<td></td>
</tr>
<tr>
<td>Liquid in a liquid</td>
<td>Vinegar</td>
<td>HC₂H₃O₂(l)</td>
<td>H₂O(l)</td>
</tr>
<tr>
<td>Solid in a liquid</td>
<td>Seawater</td>
<td>NaCl(s)</td>
<td>H₂O(l)</td>
</tr>
<tr>
<td>Tincture of iodine</td>
<td>I₂(s)</td>
<td>C₂H₅OH(l)</td>
<td></td>
</tr>
<tr>
<td><strong>Solid Solutions</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Solid in a solid</td>
<td>Brass</td>
<td>Zn(s)</td>
<td>Cu(s)</td>
</tr>
<tr>
<td>Steel</td>
<td></td>
<td>C(s)</td>
<td>Fe(s)</td>
</tr>
</tbody>
</table>

Water as a Solvent

Water is one of the most common solvents in nature. In the H₂O molecule, an oxygen atom shares electrons with two hydrogen atoms. Because oxygen is much more electronegative than hydrogen, the O—H bonds are polar. In each polar bond, the oxygen atom has a partial negative (δ⁻) charge and the hydrogen atom has a partial positive (δ⁺) charge. Because the shape of a water molecule is bent, not linear, its dipoles do not cancel out. Thus, water is polar, and is a **polar solvent**.

Attractive forces known as **hydrogen bonds** occur between molecules where partially positive hydrogen atoms are attracted to the partially negative atoms N, O, or F. As seen in the diagram, the hydrogen bonds are shown as a series of dots. Although hydrogen bonds are much weaker than covalent or ionic bonds, there are many of them linking water molecules together. Hydrogen bonds are important in the properties of biological compounds such as proteins, carbohydrates, and DNA.

In water, hydrogen bonds form between an oxygen atom in one water molecule and the hydrogen atom in another.
CHEMISTRY LINK TO HEALTH
Water in the Body

The average adult is about 60% water by mass, and the average infant about 75%. About 60% of the body’s water is contained within the cells as intracellular fluids; the other 40% makes up extracellular fluids, which include the interstitial fluid in tissue and the plasma in the blood. These external fluids carry nutrients and waste materials between the cells and the circulatory system.

Typical water gain and loss during 24 hours

<table>
<thead>
<tr>
<th>Water Gain</th>
<th>Water Loss</th>
</tr>
</thead>
<tbody>
<tr>
<td>Liquid</td>
<td>Urine</td>
</tr>
<tr>
<td>1000 mL</td>
<td>1500 mL</td>
</tr>
<tr>
<td>Food</td>
<td>Perspiration</td>
</tr>
<tr>
<td>1200 mL</td>
<td>300 mL</td>
</tr>
<tr>
<td>Metabolism</td>
<td>Breath</td>
</tr>
<tr>
<td>300 mL</td>
<td>600 mL</td>
</tr>
<tr>
<td>Total</td>
<td>Feces</td>
</tr>
<tr>
<td>2500 mL</td>
<td>100 mL</td>
</tr>
<tr>
<td></td>
<td>Total</td>
</tr>
<tr>
<td></td>
<td>2500 mL</td>
</tr>
</tbody>
</table>

The water lost from the body is replaced by the intake of fluids.

Every day you lose between 1500 and 3000 mL of water from the kidneys as urine, from the skin as perspiration, from the lungs as you exhale, and from the gastrointestinal tract. Serious dehydration can occur in an adult if there is a 10% net loss in total body fluid; a 20% loss of fluid can be fatal. An infant suffers severe dehydration with only a 5 to 10% loss in body fluid.

Water loss is continually replaced by the liquids and foods in the diet and from metabolic processes that produce water in the cells of the body. TABLE 12.2 lists the percentage by mass of water contained in some foods.

TABLE 12.2 Percentage of Water in Some Foods

<table>
<thead>
<tr>
<th>Food</th>
<th>Water (% by mass)</th>
<th>Food</th>
<th>Water (% by mass)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Vegetables</td>
<td>Meats/Fish</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Carrot</td>
<td>88</td>
<td>Chicken, cooked</td>
<td>71</td>
</tr>
<tr>
<td>Celery</td>
<td>94</td>
<td>Hamburger, broiled</td>
<td>60</td>
</tr>
<tr>
<td>Cucumber</td>
<td>96</td>
<td>Salmon</td>
<td>71</td>
</tr>
<tr>
<td>Tomato</td>
<td>94</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Fruits</td>
<td>Milk Products</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Apple</td>
<td>85</td>
<td>Cottage cheese</td>
<td>78</td>
</tr>
<tr>
<td>Cantaloupe</td>
<td>91</td>
<td>Milk, whole</td>
<td>87</td>
</tr>
<tr>
<td>Orange</td>
<td>86</td>
<td>Yogurt</td>
<td>88</td>
</tr>
<tr>
<td>Strawberry</td>
<td>90</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Watermelon</td>
<td>93</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Formation of Solutions

The interactions between solute and solvent will determine whether a solution will form. Initially, energy is needed to separate the particles in the solute and the solvent particles. Then energy is released as solute particles move between the solvent particles to form a solution. However, there must be attractions between the solute and the solvent particles to provide the energy for the initial separation. These attractions occur when the solute and the solvent have similar polarities. The expression “like dissolves like” is a way of saying that the polarities of a solute and a solvent must be similar in order for a solution to form (see FIGURE 12.2). In the absence of attractions between a solute and a solvent, there is insufficient energy to form a solution (see TABLE 12.3).

TABLE 12.3 Possible Combinations of Solutes and Solvents

<table>
<thead>
<tr>
<th>Solute Will Form</th>
<th>Solutions Will Not Form</th>
</tr>
</thead>
<tbody>
<tr>
<td>Polar</td>
<td>Solvent</td>
</tr>
<tr>
<td>Nonpolar</td>
<td>Solvent</td>
</tr>
<tr>
<td></td>
<td>Polar</td>
</tr>
<tr>
<td></td>
<td>Nonpolar</td>
</tr>
</tbody>
</table>

FIGURE 12.2 ► Like dissolves like. In each test tube, the lower layer is CH₂Cl₂ (more dense), and the upper layer is water (less dense). (a) CH₂Cl₂ is nonpolar and water is polar; the two layers do not mix. (b) The nonpolar solute I₂ (purple) is soluble in the nonpolar solvent CH₂Cl₂. (c) The ionic solute Ni(NO₃)₂ (green) is soluble in the polar solvent water.

In which layer would polar molecules of sucrose \((C_{12}H_{22}O_{11})\) be soluble?
In ionic solutes such as sodium chloride, NaCl, there are strong ionic bonds between positively charged Na\(^+\) ions and negatively charged Cl\(^-\) ions. In water, a polar solvent, the hydrogen bonds provide strong solvent–solvent attractions. When NaCl crystals are placed in water, partially negative oxygen atoms in water molecules attract positive Na\(^+\) ions, and the partially positive hydrogen atoms in other water molecules attract negative Cl\(^-\) ions (see FIGURE 12.3). As soon as the Na\(^+\) ions and the Cl\(^-\) ions form a solution, they undergo hydration as water molecules surround each ion. Hydration of the ions diminishes their attraction to other ions and keeps them in solution.

In the equation for the formation of the NaCl solution, the solid and aqueous NaCl are shown with the formula H\(_2\)O over the arrow, which indicates that water is needed for the dissociation process but is not a reactant.

\[ \text{NaCl}(s) + \text{H}_2\text{O} \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) \]

In another example, we find that a polar molecular compound such as methanol, CH\(_3\) — OH, is soluble in water because methanol has a polar —OH group that forms hydrogen bonds with water (see FIGURE 12.4). Polar solutes require polar solvents for a solution to form.

Why do KCl(s) form a solution with water, but nonpolar hexane (C\(_6\)H\(_{14}\)) does not form a solution with water?

### Solutions with Nonpolar Solutes

Compounds containing nonpolar molecules, such as iodine (I\(_2\)), oil, or grease, do not dissolve in water because there are no attractions between the particles of a nonpolar solute and the polar solvent. Nonpolar solutes require nonpolar solvents for a solution to form.

**Questions and Problems**

**12.1 Solutions**

**Learning Goal** Identify the solute and solvent in a solution; describe the formation of a solution.

12.1 Identify the solute and the solvent in each solution composed of the following:
   a. 10.0 g of NaCl and 100.0 g of H\(_2\)O
   b. 50.0 mL of ethanol, C\(_2\)H\(_5\)OH, and 10.0 mL of H\(_2\)O
   c. 0.20 L of O\(_2\) and 0.80 L of N\(_2\)

12.2 Identify the solute and the solvent in each solution composed of the following:
   a. 10.0 mL of acetic acid and 200.0 mL of water
   b. 100.0 mL of water and 5.0 g of sugar
   c. 1.0 g of Br\(_2\) and 50.0 mL of methylene chloride(l)

12.3 Describe the formation of an aqueous KI solution, when solid KI dissolves in water.

12.4 Describe the formation of an aqueous LiBr solution, when solid LiBr dissolves in water.

**Applications**

12.5 Water is a polar solvent and carbon tetrachloride (CCl\(_4\)) is a nonpolar solvent. In which solvent is each of the following, which is found or used in the body, more likely to be soluble?
   a. CaCO\(_3\) (calcium supplement), ionic
   b. retinol (vitamin A), nonpolar
   c. sucrose (table sugar), polar
   d. cholesterol (lipid), nonpolar

12.6 Water is a polar solvent and hexane (C\(_6\)H\(_{14}\)) is a nonpolar solvent. In which solvent is each of the following, which is found or used in the body, more likely to be soluble?
   a. vegetable oil, nonpolar
   b. oleic acid (lipid), nonpolar
   c. niacin (vitamin B\(_3\)), polar
   d. FeSO\(_4\) (iron supplement), ionic
12.2 Electrolytes and Nonelectrolytes

**LEARNING GOAL** Identify solutes as electrolytes or nonelectrolytes.

Solutions can be classified by their ability to conduct an electrical current. When **electrolytes** dissolve in water, the process of **dissociation** separates them into ions forming solutions that conduct electricity. When **nonelectrolytes** dissolve in water, they do not separate into ions and their solutions do not conduct electricity.

To test solutions for the presence of ions, we can use an apparatus that consists of a battery and a pair of electrodes connected by wires to a light bulb. The light bulb glows when electricity can flow, which can only happen when electrolytes provide ions that move between the electrodes to complete the circuit.

**Types of Electrolytes**

Electrolytes can be further classified as **strong electrolytes** or **weak electrolytes**. For a **strong electrolyte**, such as sodium chloride (NaCl), there is 100% dissociation of the solute into ions. When the electrodes from the light bulb apparatus are placed in the NaCl solution, the light bulb glows very brightly.

In an equation for dissociation of a compound in water, the charges must balance. For example, magnesium nitrate dissociates to give one magnesium ion for every two nitrate ions. However, only the ionic bonds between Mg$^{2+}$ and NO$_3^-$ are broken, not the covalent bonds within the polyatomic ion. The equation for the dissociation of Mg(NO$_3$)$_2$ is written as follows:

$$\text{Mg(NO}_3\text{)}_2(s) \xrightarrow{\text{Dissociation}} \text{Mg}^{2+}(aq) + 2\text{NO}_3^-(aq)$$

A **weak electrolyte** is a compound that dissolves in water mostly as molecules. Only a few of the dissolved solute molecules undergo dissociation, producing a small number of ions in solution. Thus, solutions of weak electrolytes do not conduct electrical current as well as solutions of strong electrolytes. When the electrodes are placed in a solution of a weak electrolyte, the glow of the light bulb is very dim. In an aqueous solution of the weak electrolyte HF, a few HF molecules dissociate to produce H$^+$ and F$^-$ ions. As more H$^+$ and F$^-$ ions form, some recombine to give HF molecules. These forward and reverse reactions of molecules to ions and back again are indicated by two arrows between reactant and products that point in opposite directions:

$$\text{HF}(aq) \xrightarrow{\text{Dissociation}} \text{H}^+(aq) + \text{F}^-(aq) \xleftarrow{\text{Recombination}}$$

A nonelectrolyte such as methanol (CH$_3$OH) dissolves in water only as molecules, which do not ionize. When electrodes of the light bulb apparatus are placed in a solution of a nonelectrolyte, the light bulb does not glow, because the solution does not contain ions and cannot conduct electricity.

$$\text{CH}_3\text{OH}(l) \xrightarrow{\text{H}_2\text{O}} \text{CH}_3\text{OH}(aq)$$

**TABLE 12.4** summarizes the classification of solutes in aqueous solutions.

<table>
<thead>
<tr>
<th><strong>Type of Solute</strong></th>
<th><strong>In Solution</strong></th>
<th><strong>Type(s) of Particles in Solution</strong></th>
<th><strong>Conducts Electricity?</strong></th>
<th><strong>Examples</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>Strong electrolyte</td>
<td>Dissociates completely</td>
<td>Only ions</td>
<td>Yes</td>
<td>Ionic compounds such as NaCl, KBr, MgCl$_2$, NaNO$_3$; bases such as NaOH, KOH; acids such as HCl, HBr, HI, HNO$_3$, HClO$_4$, H$_2$SO$_4$</td>
</tr>
<tr>
<td>Weak electrolyte</td>
<td>Ionizes partially</td>
<td>Mostly molecules and a few ions</td>
<td>Weakly</td>
<td>HF, H$_2$O, NH$_3$, HC$_2$H$_3$O$_2$ (acetic acid)</td>
</tr>
<tr>
<td>Nonelectrolyte</td>
<td>No ionization</td>
<td>Only molecules</td>
<td>No</td>
<td>Carbon compounds such as CH$_3$OH (methanol), C$_2$H$<em>5$OH (ethanol), C$</em>{12}$H$_2$O$_11$ (sucrose), CH$_3$N$_2$O (urea)</td>
</tr>
</tbody>
</table>
SAMPLE PROBLEM 12.1 Solutions of Electrolytes and Nonelectrolytes

Indicate whether solutions of each of the following contain only ions, only molecules, or mostly molecules and a few ions. Write the equation for the formation of a solution for each of the following:

a. Na₂SO₄(s), a strong electrolyte
b. sucrose, C₁₂H₂₂O₁₁(s), a nonelectrolyte
c. acetic acid, HC₂H₃O₂(l), a weak electrolyte

TRY IT FIRST

SOLUTION

a. An aqueous solution of Na₂SO₄(s) contains only the ions Na⁺ and SO₄²⁻.

\[
\text{Na}_2\text{SO}_4(s) \xrightarrow{\text{H}_2\text{O}} 2\text{Na}^+(aq) + \text{SO}_4^{2-}(aq)
\]

b. A nonelectrolyte such as sucrose, C₁₂H₂₂O₁₁(s), produces only molecules when it dissolves in water.

\[
\text{C}_{12}\text{H}_{22}\text{O}_{11}(s) \xrightarrow{\text{H}_2\text{O}} \text{C}_{12}\text{H}_{22}\text{O}_{11}(aq)
\]

c. A weak electrolyte such as HC₂H₃O₂(l) produces mostly molecules and a few ions when it dissolves in water.

\[
\text{HC}_2\text{H}_3\text{O}_2(l) \xrightarrow{\text{H}_2\text{O}} \text{H}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq)
\]

STUDY CHECK 12.1

Boric acid, H₃BO₃(s), is a weak electrolyte. Would you expect a boric acid solution to contain only ions, only molecules, or mostly molecules and a few ions?

ANSWER

A solution of a weak electrolyte would contain mostly molecules and a few ions.

CHEMISTRY LINK TO HEALTH

Electrolytes in Body Fluids

Electrolytes in the body play an important role in maintaining the proper function of the cells and organs in the body. Typically, the electrolytes sodium, potassium, chloride, and bicarbonate are measured in a blood test. Sodium ions regulate the water content in the body and are important in carrying electrical impulses through the nervous system. Potassium ions are also involved in the transmission of electrical impulses and play a role in the maintenance of a regular heartbeat. Chloride ions control the balance of fluids in the body. Bicarbonate is important in maintaining the proper pH of the blood. There is a charge balance because the total number of positive charges is equal to the total number of negative charges. Sometimes when vomiting, diarrhea, or sweating is excessive, the concentrations of certain electrolytes may decrease. Then fluids such as Pedialyte may be given to return electrolyte levels to normal.
The term **solubility** is used to describe the amount of a solute that can dissolve in a given amount of solvent. Many factors, such as the type of solute, the type of solvent, and the temperature, affect the solubility of a solute. **Solubility**, usually expressed in grams of solute in 100. g of solvent, is the maximum amount of solute that can be dissolved at a certain temperature. If a solute readily dissolves when added to the solvent, the solution does not contain the maximum amount of solute. We call this solution an **unsaturated solution**.

A solution that contains all the solute that can dissolve is a **saturated solution**. When a solution is saturated, the rate at which the solute dissolves becomes equal to the rate at which solid forms, a process known as **recrystallization**. Then there is no further change in the amount of dissolved solute in solution.

\[
\text{Solute + solvent} \quad \xleftrightarrow{\text{Solute dissolves}} \quad \text{saturated solution} \xrightarrow{\text{Solute recrystallizes}}
\]

We can prepare a saturated solution by adding an amount of solute greater than that needed to reach solubility. Stirring the solution will dissolve the maximum amount of solute and leave the excess on the bottom of the container. Once we have a saturated solution, the addition of more solute will only increase the amount of undissolved solute.
SAMPLE PROBLEM 12.2 Saturated Solutions

At 20 °C, the solubility of KCl is 34 g/100 g of H₂O. In the laboratory, a student mixes 75 g of KCl with 200 g of H₂O at a temperature of 20 °C.

a. How much of the KCl will dissolve?

b. Is the solution saturated or unsaturated?

c. What is the mass, in grams, of any solid KCl left undissolved on the bottom of the container?

SOLUTION

a. At 20 °C, KCl has a solubility of 34 g of KCl in 100 g of water. Using the solubility as a conversion factor, we can calculate the maximum amount of KCl that can dissolve in 200 g of water as follows:

\[
200 \text{ g H}_2\text{O} \times \frac{34 \text{ g KCl}}{100 \text{ g H}_2\text{O}} = 68 \text{ g of KCl}
\]

b. Because 75 g of KCl exceeds the maximum amount (68 g) that can dissolve in 200 g of water, the KCl solution is saturated.

c. If we add 75 g of KCl to 200 g of water and only 68 g of KCl can dissolve, there is 7 g (75 g - 68 g) of solid (undissolved) KCl on the bottom of the container.

STUDY CHECK 12.2

At 40 °C, the solubility of KNO₃ is 65 g/100 g of H₂O. How many grams of KNO₃ will dissolve in 120 g of H₂O at 40 °C?

ANSWER

78 g of KNO₃
Kidney stones are solid materials that form in the urinary tract. Most kidney stones are composed of calcium phosphate and calcium oxalate, although they can be solid uric acid. Insufficient water intake and high levels of calcium, oxalate, and phosphate in the urine can lead to the formation of kidney stones. When a kidney stone passes through the urinary tract, it causes considerable pain and discomfort, necessitating the use of painkillers and surgery. Sometimes ultrasound is used to break up kidney stones. Persons prone to kidney stones are advised to drink six to eight glasses of water every day to prevent saturation levels of minerals in the urine.

Kidney stones form when calcium phosphate exceeds its solubility.

**Effect of Temperature on Solubility**

The solubility of most solids is greater as temperature increases, which means that solutions usually contain more dissolved solute at higher temperature. A few substances show little change in solubility at higher temperatures, and a few are less soluble (see [FIGURE 12.5](#)). For example, when you add sugar to iced tea, some undissolved sugar may form on the bottom of the glass. But if you add sugar to hot tea, many teaspoons of sugar are needed before solid sugar appears. Hot tea dissolves more sugar than does cold tea because the solubility of sugar is much greater at a higher temperature.

When a saturated solution is carefully cooled, it becomes a *supersaturated solution* because it contains more solute than the solubility allows. Such a solution is unstable, and if the solution is agitated or if a solute crystal is added, the excess solute will recrystallize to give a saturated solution again.

![FIGURE 12.5](#) In water, most common solids are more soluble as the temperature increases.

**Compare the solubility of NaNO₃ at 20 °C and 60 °C.**

Conversely, the solubility of a gas in water decreases as the temperature increases. At higher temperatures, more gas molecules have the energy to escape from the solution. Perhaps you have observed the bubbles escaping from a cold carbonated soft drink as it warms. At high temperatures, bottles containing carbonated solutions may burst as more gas molecules leave the solution and increase the gas pressure inside the bottle. Biologists have found that increased temperatures in rivers and lakes cause the amount of dissolved oxygen to decrease until the warm water can no longer support a biological community. Electricity-generating plants are required to have their own ponds to use with their cooling towers to lessen the threat of thermal pollution in surrounding waterways.

![In water, gases are less soluble as the temperature increases.](#)
Henry’s Law

Henry’s law states that the solubility of gas in a liquid is directly related to the pressure of that gas above the liquid. At higher pressures, there are more gas molecules available to enter and dissolve in the liquid. A can of soda is carbonated by using CO₂ gas at high pressure to increase the solubility of the CO₂ in the beverage. When you open the can at atmospheric pressure, the pressure on the CO₂ drops, which decreases the solubility of CO₂. As a result, bubbles of CO₂ rapidly escape from the solution. The burst of bubbles is even more noticeable when you open a warm can of soda.

Soluble and Insoluble Ionic Compounds

In our discussion up to now, we have considered ionic compounds that dissolve in water. However, some ionic compounds do not dissociate into ions and remain as solids even in contact with water. The solubility rules give some guidelines about the solubility of ionic compounds in water.

Ionic compounds that are soluble in water typically contain at least one of the ions in Table 12.5. Only an ionic compound containing a soluble cation or anion will dissolve in water. Most ionic compounds containing Cl⁻ are soluble, but AgCl, PbCl₂, and Hg₂Cl₂ are insoluble. Similarly, most ionic compounds containing SO₄²⁻ are soluble, but a few are insoluble. Most other ionic compounds are insoluble (see Figure 12.6). In an insoluble ionic compound, the ionic bonds between its positive and negative ions are too strong for the polar water molecules to break. We can use the solubility rules to predict whether a solid ionic compound would be soluble or not. Table 12.6 illustrates the use of these rules.

<table>
<thead>
<tr>
<th>TABLE 12.5 Solubility Rules for Ionic Compounds in Water</th>
</tr>
</thead>
<tbody>
<tr>
<td>An ionic compound is soluble in water if it contains one of the following:</td>
</tr>
<tr>
<td>Positive Ions:</td>
</tr>
<tr>
<td>Negative Ions:</td>
</tr>
<tr>
<td>Cl⁻, Br⁻, I⁻ except when combined with Ag⁺, Pb²⁺, or Hg₂²⁺</td>
</tr>
<tr>
<td>SO₄²⁻ except when combined with Ba²⁺, Pb²⁺, Ca²⁺, Sr²⁺, or Hg₂²⁺</td>
</tr>
<tr>
<td>Ionic compounds that do not contain at least one of these ions are usually insoluble.</td>
</tr>
</tbody>
</table>
## 12.3 Solubility

**Using Solubility Rules**

<table>
<thead>
<tr>
<th>Ionic Compound</th>
<th>Solubility in Water</th>
<th>Reasoning</th>
</tr>
</thead>
<tbody>
<tr>
<td>K$_2$S</td>
<td>Soluble</td>
<td>Contains K$^+$</td>
</tr>
<tr>
<td>Ca(NO$_3$)$_2$</td>
<td>Soluble</td>
<td>Contains NO$_3^-$</td>
</tr>
<tr>
<td>PbCl$_2$</td>
<td>Insoluble</td>
<td>Is an insoluble chloride</td>
</tr>
<tr>
<td>NaOH</td>
<td>Soluble</td>
<td>Contains Na$^+$</td>
</tr>
<tr>
<td>AlPO$_4$</td>
<td>Insoluble</td>
<td>Contains no soluble ions</td>
</tr>
</tbody>
</table>

**In medicine, insoluble BaSO$_4$ is used as an opaque substance to enhance X-rays of the gastrointestinal tract (see FIGURE 12.7). BaSO$_4$ is so insoluble that it does not dissolve in gastric fluids. Other ionic barium compounds cannot be used because they would dissolve in water, releasing Ba$^{2+}$ which is poisonous.**

**SAMPLE PROBLEM 12.3 Soluble and Insoluble Ionic Compounds**

Predict whether each of the following ionic compounds is soluble in water and explain why:

a. Na$_3$PO$_4$

b. CaCO$_3$

**SOLUTION**

a. The ionic compound Na$_3$PO$_4$ is soluble in water because any compound that contains Na$^+$ is soluble.

b. The ionic compound CaCO$_3$ is not soluble. The compound does not contain a soluble positive ion, which means that ionic compound containing Ca$^{2+}$ and CO$_3^{2-}$ is not soluble.

**STUDY CHECK 12.3**

In some electrolyte drinks, MgCl$_2$ is added to provide magnesium. Why would you expect MgCl$_2$ to be soluble in water?

**ANSWER**

MgCl$_2$ is soluble in water because ionic compounds that contain chloride are soluble unless they contain Ag$^+$, Pb$^{2+}$, or Hg$_2^{2+}$.

**Formation of a Solid**

We can use solubility rules to predict whether a solid, called a *precipitate*, forms when two solutions containing soluble reactants are mixed as shown in Sample Problem 12.4.
SAMPLE PROBLEM 12.4 Writing Equations for the Formation of an Insoluble Ionic Compound

When solutions of NaCl and AgNO₃ are mixed, a white solid forms. Write the ionic and net ionic equations for the reaction.

TRY IT FIRST

SOLUTION

STEP 1 Write the ions of the reactants.

<table>
<thead>
<tr>
<th>Reactants (initial combinations)</th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Ag⁺(aq) + NO₃⁻(aq)</td>
<td>Na⁺(aq) + Cl⁻(aq)</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Write the combinations of ions and determine if any are insoluble.

Mixture (new combinations) | Product | Soluble |
--- | --- | ---|
Ag⁺(aq) + Cl⁻(aq) | AgCl | No |
Na⁺(aq) + NO₃⁻(aq) | NaNO₃ | Yes |

STEP 3 Write the ionic equation including any solid. In the ionic equation, we show all the ions of the reactants. The products include the solid AgCl that forms.

Ag⁺(aq) + NO₃⁻(aq) + Na⁺(aq) + Cl⁻(aq) → AgCl(s) + Na⁺(aq) + NO₃⁻(aq)

STEP 4 Write the net ionic equation. To write a net ionic equation, we remove the Na⁺ and NO₃⁻ ions, known as spectator ions, which are unchanged.

Ag⁺(aq) + Cl⁻(aq) → AgCl(s)

ENGAGE

When mixing solutions of Pb(NO₃)₂ and NaBr, how do we know that PbBr₂(s) is the solid that forms?
12.3 Solubility

Study Check 12.4
Predict whether a solid might form in each of the following mixtures of solutions. If so, write the net ionic equation for the reaction.

a. $\text{NH}_4\text{Cl}(aq) + \text{Ca(NO}_3\text{)}_2(aq)$  
   b. $\text{Pb(NO}_3\text{)}_2(aq) + \text{KCl}(aq)$

Answer
a. No solid forms because the products, $\text{NH}_4\text{NO}_3(aq)$ and $\text{CaCl}_2(aq)$, are soluble.
     
   b. $\text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{PbCl}_2(s)$

Questions and Problems

12.3 Solubility

Learning Goal Define solubility; distinguish between an unsaturated and a saturated solution. Identify an ionic compound as soluble or insoluble.

12.15 State whether each of the following refers to a saturated or an unsaturated solution:
   a. A crystal added to a solution does not change in size.
   b. A sugar cube completely dissolves when added to a cup of coffee.

12.16 State whether each of the following refers to a saturated or an unsaturated solution:
   a. A spoonful of salt added to boiling water dissolves.
   b. A layer of sugar forms on the bottom of a glass of tea as ice is added.

Use the following table for problems 12.17 to 12.20:

<table>
<thead>
<tr>
<th>Solubility (g/100. g H$_2$O)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Substance</td>
</tr>
<tr>
<td>KCl</td>
</tr>
<tr>
<td>NaNO$_3$</td>
</tr>
<tr>
<td>C$<em>{12}$H$</em>{22}$O$_{11}$ (sugar)</td>
</tr>
</tbody>
</table>

12.17 Determine whether each of the following solutions will be saturated or unsaturated at 20 °C:
   a. adding 25 g of KCl to 100. g of H$_2$O
   b. adding 11 g of NaNO$_3$ to 25 g of H$_2$O
   c. adding 400. g of sugar to 125 g of H$_2$O

12.18 Determine whether each of the following solutions will be saturated or unsaturated at 50 °C:
   a. adding 25 g of KCl to 50. g of H$_2$O
   b. adding 150. g of NaNO$_3$ to 75 g of H$_2$O
   c. adding 80. g of sugar to 25 g of H$_2$O

12.19 A solution containing 80. g of KCl in 200. g of H$_2$O at 50 °C is cooled to 20 °C:
   a. How many grams of KCl remain in solution at 20 °C?
   b. How many grams of solid KCl crystallized after cooling?

12.20 A solution containing 80. g of NaNO$_3$ in 75 g of H$_2$O at 50 °C is cooled to 20 °C.
   a. How many grams of NaNO$_3$ remain in solution at 20 °C?
   b. How many grams of solid NaNO$_3$ crystallized after cooling?

12.21 Explain the following observations:
   a. More sugar dissolves in hot tea than in iced tea.
   b. Champagne in a warm room goes flat.
   c. A warm can of soda has more spray when opened than a cold one.

12.22 Explain the following observations:
   a. An open can of soda loses its “fizz” faster at room temperature than in the refrigerator.
   b. Chlorine gas in tap water escapes as the sample warms to room temperature.
   c. Less sugar dissolves in iced coffee than in hot coffee.

12.23 Predict whether each of the following ionic compounds is soluble in water:
   a. LiCl
   b. PbS
   c. BaCO$_3$
   d. K$_2$O
   e. Fe(NO$_3$)$_3$

12.24 Predict whether each of the following ionic compounds is soluble in water:
   a. AgCl
   b. KI
   c. Na$_2$S
   d. Ag$_2$O
   e. CaSO$_4$

12.25 Determine whether a solid forms when solutions containing the following ionic compounds are mixed. If so, write the ionic equation and the net ionic equation.
   a. KCl($aq$) and Na$_2$S($aq$)
   b. $\text{AgNO}_3(aq)$ and K$_2$S($aq$)
   c. CaCl$_2(aq)$ and Na$_2$SO$_4(aq)$
   d. CuCl$_2(aq)$ and Li$_3$PO$_4(aq)$

12.26 Determine whether a solid forms when solutions containing the following ionic compounds are mixed. If so, write the ionic equation and the net ionic equation.
   a. Na$_3$PO$_4(aq)$ and AgNO$_3(aq)$
   b. K$_2$SO$_4(aq)$ and Na$_2$CO$_3(aq)$
   c. Pb(NO$_3$)$_2(aq)$ and Na$_2$CO$_3(aq)$
   d. BaCl$_2(aq)$ and KOH($aq$)
12.4 Solution Concentrations

LEARNING GOAL. Calculate the concentration of a solute in a solution; use concentration as conversion factors to calculate the amount of solute or solution.

Our body fluids contain water and dissolved substances including glucose, urea, and electrolytes such as \( \text{K}^+ \), \( \text{Na}^+ \), \( \text{Cl}^- \), \( \text{Mg}^{2+} \), \( \text{HCO}_3^- \), and \( \text{HPO}_4^{2-} \). Proper amounts of each of these dissolved substances and water must be maintained in the body fluids. Small changes in electrolyte levels can seriously disrupt cellular processes and endanger our health. Solutions can be described by their concentration, which is the amount of solute in a specific amount of that solution as shown in TABLE 12.7. The amount of solute dissolved in a certain amount of solution is called the concentration of the solution. We will look at ways to express a concentration as a ratio of a certain amount of solute in a given amount of solution. The amount of a solute may be expressed in units of grams, milliliters, or moles. The amount of a solution may be expressed in units of grams, milliliters, or liters.

Concentration of a solution = \[ \frac{\text{amount of solute}}{\text{amount of solution}} \]

| TABLE 12.7 Summary of Types of Concentration Expressions and Their Units |
|----------------------------------|---------------------------------|-----------------|------------------|------------------|
| Concentration                   | Mass Percent (m/m) | Volume Percent (v/v) | Mass/Volume Percent (m/v) | Molarity (M) |
| Units                            | g                  | mL               | g                | mole          |
| Solute                           | g                  | mL               | g                | mole          |
| Solution                         | g                  | mL               | mL               | L             |

Mass Percent (m/m) Concentration

Mass percent (m/m) describes the mass of the solute in grams for 100. g of solution. The mass percent is calculated by dividing the mass of a solute by the mass of the solution multiplied by 100% to give the percentage. In the calculation of mass percent (m/m), the units of mass of the solute and solution must be the same. If the mass of the solute is given as grams, then the mass of the solution must also be grams. The mass of the solution is the sum of the mass of the solute and the mass of the solvent.

\[
\text{Mass percent (m/m)} = \frac{\text{mass of solute (g)}}{\text{mass of solute (g)} + \text{mass of solvent (g)}} \times 100\%
\]

Suppose we prepared a solution by mixing 8.00 g of KCl (solute) with 42.00 g of water (solvent). Together, the mass of the solute and mass of solvent give the mass of the solution (8.00 g + 42.00 g = 50.00 g). Mass percent is calculated by substituting the mass of the solute and the mass of the solution into the mass percent expression.

\[
\frac{8.00 \text{ g KCl}}{50.00 \text{ g solution}} \times 100\% = 16.0\% \text{ (m/m) KCl solution}
\]

\[
\frac{8.00 \text{ g KCl} + 42.00 \text{ g H}_2\text{O}}{\text{Solute} + \text{Solvent}}
\]
SAMPLE PROBLEM 12.5 Calculating Mass Percent (m/m) Concentration

What is the mass percent of NaOH in a solution prepared by dissolving 30.0 g of NaOH in 120.0 g of H₂O?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>30.0 g of NaOH, 120.0 g of H₂O</td>
<td>mass percent (m/m)</td>
<td>mass of solute ÷ mass of solution × 100%</td>
</tr>
</tbody>
</table>

STEP 2 Write the concentration expression.

Mass percent (m/m) = \( \frac{\text{grams of solute}}{\text{grams of solution}} \times 100\% \)

STEP 3 Substitute solute and solution quantities into the expression and calculate. The mass of the solution is obtained by adding the mass of the solute and the mass of the solution.

\[
\text{mass of solution} = 30.0 \text{ g NaOH} + 120.0 \text{ g H}_2\text{O} = 150.0 \text{ g of NaOH solution}
\]

Three SFs

\[
\text{Mass percent (m/m)} = \frac{30.0 \text{ g NaOH}}{150.0 \text{ g solution}} \times 100\% 
\]

Four SFs

\[
= 20.0\% \text{ (m/m) NaOH solution}
\]

Three SFs

STUDY CHECK 12.5

What is the mass percent (m/m) of NaCl in a solution made by dissolving 2.0 g of NaCl in 56.0 g of H₂O?

ANSWER

3.4% (m/m) NaCl solution

Using Mass Percent Concentration as a Conversion Factor

In the preparation of solutions, we often need to calculate the amount of solute or solution. Then the concentration of a solution is useful as a conversion factor as shown in Sample Problem 12.6.

SAMPLE PROBLEM 12.6 Using Mass Percent to Find Mass of Solute

The topical antibiotic ointment Neosporin is 3.5% (m/m) neomycin solution. How many grams of neomycin are in a tube containing 64 g of ointment?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>64 g of 3.5% (m/m) neomycin solution</td>
<td>grams of neomycin</td>
<td>mass percent factor ( \frac{\text{g of solute}}{100, \text{ g of solution}} )</td>
</tr>
</tbody>
</table>
Guide to Using Concentration to Calculate Mass or Volume

STEP 1 State the given and needed quantities.

STEP 2 Write a plan to calculate the mass or volume.

STEP 3 Write equalities and conversion factors. The mass percent (m/m) indicates the grams of a solute in every 100. g of a solution. The mass percent (3.5% m/m) can be written as two conversion factors.

\[
\frac{3.5 \text{ g of neomycin}}{100. \text{ g ointment}} = \frac{3.5 \text{ g of neomycin}}{100. \text{ g ointment}} \quad \text{and} \quad \frac{3.5 \text{ g of neomycin}}{3.5 \text{ g neomycin}}
\]

STEP 4 Set up the problem to calculate the mass.

\[
64 \text{ g ointment} \times \frac{3.5 \text{ g of neomycin}}{100. \text{ g ointment}} = 2.2 \text{ g of neomycin}
\]

EnGAGE

How is the mass percent (m/m) of a solution used to convert the mass of the solution to the grams of solute?

Volume Percent (v/v) Concentration

Because the volumes of liquids or gases are easily measured, the concentrations of their solutions are often expressed as volume percent (v/v). The units of volume used in the ratio must be the same, for example, both in milliliters or both in liters.

\[
\text{Volume percent (v/v) } = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%
\]

We interpret a volume percent as the volume of solute in 100. mL of solution. On a bottle of extract of vanilla, a label that reads alcohol 35% (v/v) means 35 mL of ethanol solute in 100. mL of vanilla solution.

SAMPLE PROBLEM 12.7 Calculating Volume Percent (v/v) Concentration

A bottle contains 59 mL of lemon extract solution. If the extract contains 49 mL of alcohol, what is the volume percent (v/v) of the alcohol in the solution?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

\[
\begin{array}{ccc}
\text{ANALYZE THE PROBLEM} & \text{Given} & \text{Need} & \text{Connect} \\
49 \text{ mL of alcohol} & \text{volume} & \text{volume of solute} & \text{volume of solution} \\
59 \text{ mL of solution} & \text{percent (v/v)} & \text{volume of solution} & \times 100\%
\end{array}
\]

STEP 2 Write the concentration expression.

\[
\text{Volume percent (v/v) } = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%
\]
STEP 3 Substitute solute and solution quantities into the expression and calculate.

Volume percent (v/v) = \( \frac{49 \text{ mL alcohol}}{59 \text{ mL solution}} \times 100\% \)

= 83% (v/v) alcohol solution

STUDY CHECK 12.7

What is the volume percent (v/v) of Br\(_2\) in a solution prepared by dissolving 12 mL of liquid bromine (Br\(_2\)) in the solvent carbon tetrachloride (CCl\(_4\)) to make 250 mL of solution?

ANSWER

4.8% (v/v) Br\(_2\) in CCl\(_4\)

Mass/Volume Percent (m/v) Concentration

Mass/volume percent (m/v) describes the mass of the solute in grams for exactly 100. mL of solution. In the calculation of mass/volume percent, the unit of mass of the solute is grams and the unit of the solution volume is milliliters.

\[
\text{Mass/volume percent (m/v)} = \frac{\text{grams of solute}}{\text{milliliters of solution}} \times 100\%
\]

The mass/volume percent is widely used in hospitals and pharmacies for the preparation of intravenous solutions and medicines. For example, a 5% (m/v) glucose solution contains 5 g of glucose in 100. mL of solution. The volume of solution represents the combined volumes of the glucose and H\(_2\)O.

SAMPLE PROBLEM 12.8 Calculating Mass/Volume Percent (m/v) Concentration

A potassium iodide solution may be used in a diet that is low in iodine. A KI solution is prepared by dissolving 5.0 g of KI in enough water to give a final volume of 250 mL. What is the mass/volume percent (m/v) of the KI solution?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>5.0 g of KI solute, 250 mL of KI solution</td>
<td>mass/volume percent (m/v)</td>
<td>mass of solute/volume of solution ( \times 100% )</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Write the concentration expression.

Mass/volume percent (m/v) = \( \frac{\text{mass of solute}}{\text{volume of solution}} \times 100\% \)

STEP 3 Substitute solute and solution quantities into the expression and calculate.

Mass/volume percent (m/v) = \( \frac{5.0 \text{ g KI}}{250 \text{ mL solution}} \times 100\% = 2.0\% \text{ (m/v) KI solution} \)
STUDY CHECK 12.8
What is the mass/volume percent (m/v) of NaOH in a solution prepared by dissolving 12 g of NaOH in enough water to make 220 mL of solution?

ANSWER
5.5% (m/v) NaOH solution

SAMPLE PROBLEM 12.9 Using Mass/Volume Percent to Find Mass of Solute
A topical antibiotic is 1.0% (m/v) clindamycin. How many grams of clindamycin are in 60 mL of the 1.0% (m/v) solution?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>60. mL of 1.0% (m/v) clindamycin solution</td>
<td>grams of clindamycin</td>
<td>% (m/v) factor</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to calculate the mass.

milliliters of solution % (m/v) factor grams of clindamycin

STEP 3 Write equalities and conversion factors. The percent (m/v) indicates the grams of a solute in every 100 mL of a solution. The 1.0% (m/v) can be written as two conversion factors.

\[
\begin{align*}
1.0 \text{ g clindamycin} &= 100. \text{ mL of solution} \\
1.0 \text{ g clindamycin} &= 100. \text{ mL solution} \\
100. \text{ mL solution} &= 1.0 \text{ g clindamycin}
\end{align*}
\]

STEP 4 Set up the problem to calculate the mass. The volume of the solution is converted to mass of solute using the conversion factor that cancels mL.

\[
60. \text{ mL solution} \times \frac{1.0 \text{ g clindamycin}}{100. \text{ mL solution}} = 0.60 \text{ g of clindamycin}
\]

STUDY CHECK 12.9
In 2010, the FDA approved a 2.0% (m/v) morphine oral solution to treat severe or chronic pain. How many grams of morphine does a patient receive if 0.60 mL of 2.0% (m/v) morphine solution was ordered?

ANSWER
0.012 g of morphine

Molarity (M) Concentration
When chemists work with solutions, they often use molarity (M), a concentration that states the number of moles of solute in exactly 1 L of solution.

\[
\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}
\]
The molarity of a solution can be calculated when we know the moles of solute and the volume of solution in liters. For example, if 1.0 mol of NaCl were dissolved in enough water to prepare 1.0 L of solution, the resulting NaCl solution has a molarity of 1.0 M. The abbreviation M indicates the units of mole per liter (mol/L).

\[ M = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{1.0 \text{ mol NaCl}}{1 \text{ L solution}} = 1.0 \text{ M NaCl solution} \]

**SAMPLE PROBLEM 12.10 Calculating Molarity**

What is the molarity (M) of 60.0 g of NaOH in 0.250 L of NaOH solution?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>60.0 g of NaOH, 0.250 L of NaOH solution</td>
<td>molarity (mol/L)</td>
<td>molar mass of NaOH, moles of solute</td>
<td>liters of solution</td>
</tr>
</tbody>
</table>

To calculate the moles of NaOH, we need to write the equality and conversion factors for the molar mass of NaOH. Then the moles in 60.0 g of NaOH can be determined.

\[ \begin{align*}
1 \text{ mol NaOH} &= \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}} \\
40.00 \text{ g NaOH} &= 1 \text{ mol NaOH} \\
1 \text{ mol NaOH} &= 40.00 \text{ g NaOH} \\
\end{align*} \]

moles of NaOH = \( 60.0 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} \)  
= 1.50 mol of NaOH  
volume of solution = 0.250 L of NaOH solution

**STEP 2** Write the concentration expression.

\[ \text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}} \]

**STEP 3** Substitute solute and solution quantities into the expression and calculate.

\[ M = \frac{1.50 \text{ mol NaOH}}{0.250 \text{ L solution}} = \frac{6.00 \text{ mol NaOH}}{1 \text{ L solution}} = 6.00 \text{ M NaOH solution} \]

**STUDY CHECK 12.10**

What is the molarity of a solution that contains 75.0 g of KNO₃ dissolved in 0.350 L of solution?

**ANSWER**

2.12 M KNO₃ solution
SAMPLE PROBLEM 12.11 Using Molarity to Calculate Volume of Solution

How many liters of a 2.00 M NaCl solution are needed to provide 67.3 g of NaCl?

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>67.3 g of NaCl, 2.00 M NaCl solution</td>
<td>liters of NaCl solution</td>
<td>molar mass of NaCl, molarity</td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to calculate the volume.

grams of NaCl \( \rightarrow \) Molar mass \( \rightarrow \) moles of NaCl \( \rightarrow \) Molarity \( \rightarrow \) liters of NaCl solution

STEP 3 Write equalities and conversion factors.

\[
\begin{align*}
1 \text{ mol of NaCl} &= \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} \\
67.3 \text{ g NaCl} &= 1.14 \text{ mol NaCl} \\
1 \text{ L NaCl solution} &= \frac{2.00 \text{ mol NaCl}}{1 \text{ L NaCl solution}} \\
0.576 \text{ L of NaCl solution} &= \frac{2.00 \text{ mol NaCl}}{1 \text{ L NaCl solution}}
\end{align*}
\]

STEP 4 Set up the problem to calculate the volume.

\[
67.3 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \times \frac{1 \text{ L NaCl solution}}{2.00 \text{ mol NaCl}} = 0.576 \text{ L of NaCl solution}
\]

STUDY CHECK 12.11

How many milliliters of a 6.0 M HCl solution will provide 164 g of HCl?

ANSWER

750 mL of HCl solution

A summary of percent concentrations and molarity, their meanings, and conversion factors are given in **TABLE 12.8**.

<table>
<thead>
<tr>
<th>Percent Concentration</th>
<th>Meaning</th>
<th>Conversion Factors</th>
</tr>
</thead>
<tbody>
<tr>
<td>10% (m/m) KCl solution</td>
<td>10 g of KCl in 100. g of KCl solution</td>
<td>( \frac{10 \text{ g KCl}}{100. \text{ g solution}} ) and ( \frac{100. \text{ g solution}}{10 \text{ g KCl}} )</td>
</tr>
<tr>
<td>12% (v/v) ethanol solution</td>
<td>12 mL of ethanol in 100. mL of ethanol solution</td>
<td>( \frac{12 \text{ mL ethanol}}{100. \text{ mL solution}} ) and ( \frac{100. \text{ mL solution}}{12 \text{ mL ethanol}} )</td>
</tr>
<tr>
<td>5% (m/v) glucose solution</td>
<td>5 g of glucose in 100. mL of glucose solution</td>
<td>( \frac{5 \text{ g glucose}}{100. \text{ mL solution}} ) and ( \frac{100. \text{ mL solution}}{5 \text{ g glucose}} )</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Molarity</th>
<th>Meaning</th>
<th>Conversion Factors</th>
</tr>
</thead>
<tbody>
<tr>
<td>6.0 M HCl solution</td>
<td>6.0 mol of HCl in 1 L of HCl solution</td>
<td>( \frac{6.0 \text{ mol HCl}}{1 \text{ L solution}} ) and ( \frac{1 \text{ L solution}}{6.0 \text{ mol HCl}} )</td>
</tr>
</tbody>
</table>
12.4 Solution Concentrations

LEARNING GOAL Calculate the concentration of a solute in a solution; use concentration as conversion factors to calculate the amount of solute or solution.

12.27 What is the difference between a 5.00% (m/m) glucose solution and a 5.00% (m/v) glucose solution?

12.28 What is the difference between a 10.0% (v/v) methanol (CH₃OH) solution and a 10.0% (m/m) methanol solution?

12.29 Calculate the mass percent (m/m) for the solute in each of the following:
   a. 25 g of KCl and 125 g of H₂O
   b. 12 g of sucrose in 225 g of tea solution
   c. 8.0 g of CaCl₂ in 80.0 g of CaCl₂ solution

12.30 Calculate the mass percent (m/m) for the solute in each of the following:
   a. 75 g of NaOH in 325 g of NaOH solution
   b. 2.0 g of KOH and 20.0 g of H₂O
   c. 48.5 g of Na₂CO₃ in 250.0 g of Na₂CO₃ solution

12.31 Calculate the mass/volume percent (m/v) for the solute in each of the following:
   a. 75 g of Na₂SO₄ in 250 mL of Na₂SO₄ solution
   b. 39 g of sucrose in 355 mL of a carbonated drink

12.32 Calculate the mass/volume percent (m/v) for the solute in each of the following:
   a. 2.50 g of LiCl in 40.0 mL of LiCl solution
   b. 7.5 g of casein in 120 mL of low-fat milk

12.33 Calculate the grams or milliliters of solute needed to prepare each of the following:
   a. 50. g of a 5.0% (m/m) KCl solution
   b. 1250 mL of a 4.0% (m/v) NH₄Cl solution
   c. 250. mL of a 10.0% (v/v) acetic acid solution

12.34 Calculate the grams or milliliters of solute needed to prepare each of the following:
   a. 150. g of a 40.0% (m/m) LiNO₃ solution
   b. 450 mL of a 2.0% (m/v) KOH solution
   c. 225 mL of a 15% (v/v) isopropyl alcohol solution

12.35 A mouthwash contains 22.5% (v/v) alcohol. If the bottle of mouthwash contains 355 mL, what is the volume, in milliliters, of alcohol?

12.36 A bottle of champagne is 11% (v/v) alcohol. If there are 750 mL of champagne in the bottle, what is the volume, in milliliters, of alcohol?

12.37 For each of the following solutions, calculate the:
   a. grams of 25% (m/m) LiNO₃ solution that contains 5.0 g of LiNO₃
   b. milliliters of 10.0% (m/v) KOH solution that contains 40.0 g of KOH
   c. milliliters of 10.0% (v/v) formic acid solution that contains 2.0 mL of formic acid

12.38 For each of the following solutions, calculate the:
   a. grams of 2.0% (m/m) NaCl solution that contains 7.50 g of NaCl
   b. milliliters of 25% (m/v) NaF solution that contains 4.0 g of NaF
   c. milliliters of 8.0% (v/v) ethanol solution that contains 20.0 mL of ethanol

12.39 Calculate the molarity of each of the following:
   a. 2.00 mol of glucose in 4.00 L of a glucose solution
   b. 4.00 g of KOH in 2.00 L of a KOH solution
   c. 5.85 g of NaCl in 400. mL of a NaCl solution

12.40 Calculate the molarity of each of the following:
   a. 0.500 mol of glucose in 0.200 L of a glucose solution
   b. 73.0 g of HCl in 2.00 L of a HCl solution
   c. 30.0 g of NaOH in 350. mL of a NaOH solution

12.41 Calculate the grams of solute needed to prepare each of the following:
   a. 2.00 L of a 1.50 M NaOH solution
   b. 4.00 L of a 0.200 M KCl solution
   c. 25.0 mL of a 6.00 M HCl solution

12.42 Calculate the grams of solute needed to prepare each of the following:
   a. 2.00 L of a 6.00 M NaOH solution
   b. 5.00 L of a 0.100 M CaCl₂ solution
   c. 175 mL of a 3.00 M NaNO₃ solution

12.43 For each of the following solutions, calculate the:
   a. liters of a 2.0 M KBr solution to obtain 3.00 mol of KBr
   b. liters of a 1.50 M NaCl solution to obtain 15.0 mol of NaCl
   c. milliliters of a 0.800 M Ca(NO₃)₂ solution to obtain 0.050 mol of Ca(NO₃)₂

12.44 For each of the following solutions, calculate the:
   a. liters of a 4.00 M KCl solution to obtain 0.100 mol of KCl
   b. liters of a 6.00 M HCl solution to obtain 5.00 mol of HCl
   c. milliliters of a 2.50 M K₂SO₄ solution to obtain 1.20 mol of K₂SO₄

12.45 Calculate the volume, in milliliters, for each of the following that provides the given amount of solute:
   a. 12.5 g of Na₂CO₃ from a 0.120 M Na₂CO₃ solution
   b. 0.850 mol of NaNO₃ from a 0.500 M NaNO₃ solution
   c. 30.0 g of LiOH from a 2.70 M LiOH solution

12.46 Calculate the volume, in liters, for each of the following that provides the given amount of solute:
   a. 5.00 mol of NaOH from a 12.0 M NaOH solution
   b. 15.0 g of Na₂SO₄ from a 4.00 M Na₂SO₄ solution
   c. 28.0 g of NaHCO₃ from a 1.50 M NaHCO₃ solution

Applications

12.47 A patient receives 100. mL of 20.0% (m/v) mannitol solution every hour.
   a. How many grams of mannitol are given in 1 h?
   b. How many grams of mannitol does the patient receive in 12 h?

12.48 A patient receives 250 mL of a 4.0% (m/v) amino acid solution twice a day.
   a. How many grams of amino acids are in 250 mL of solution?
   b. How many grams of amino acids does the patient receive in 1 day?

12.49 A patient needs 100. g of glucose in the next 12 h. How many liters of a 5% (m/v) glucose solution must be given?

12.50 A patient received 2.0 g of NaCl in 8 h. How many milliliters of a 0.90% (m/v) NaCl (saline) solution were delivered?
12.5 Dilution of Solutions

LEARNING GOAL Describe the dilution of a solution; calculate the unknown concentration or volume when a solution is diluted.

In chemistry and biology, we often prepare diluted solutions from more concentrated solutions. In a process called dilution, a solvent, usually water, is added to a solution, which increases the volume. As a result, the concentration of the solution decreases. In an everyday example, you are making a dilution when you add three cans of water to a can of concentrated orange juice.

\[ \text{1 can of orange juice concentrate} + 3 \text{ cans of water} = 4 \text{ cans of orange juice} \]

Although the addition of solvent increases the volume, the amount of solute does not change; it is the same in the concentrated solution and the diluted solution (see FIGURE 12.8).

Grams or moles of solute = grams or moles of solute

\[
\text{Concentrated solution} = \text{Diluted solution}
\]

We can write this equality in terms of the concentration, \( C \), and the volume, \( V \). The concentration, \( C \), may be percent concentration or molarity.

\[
\frac{C_1V_1}{\text{Concentrated solution}} = \frac{C_2V_2}{\text{Diluted solution}}
\]

If we are given any three of the four variables (\( C_1, C_2, V_1, \) or \( V_2 \)), we can rearrange the dilution expression to solve for the unknown quantity as seen in Sample Problem 12.12.

**FIGURE 12.8** When water is added to a concentrated solution, there is no change in the number of particles. However, the solute particles spread out as the volume of the diluted solution increases.

**SAMPLE PROBLEM 12.12 Dilution of a Solution**

A doctor orders 1000. mL of a 35.0% (m/v) dextrose solution. If you have a 50.0% (m/v) dextrose solution, how many milliliters would you use to prepare 1000. mL of 35.0% (m/v) dextrose solution?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** Prepare a table of the concentrations and volumes of the solutions. For our problem analysis, we organize the solution data in a table, making sure that the units of concentration and volume are the same.
12.5 Dilution of Solutions

Guide to Calculating Dilution Quantities

**STEP 1** Prepare a table of the concentrations and volumes of the solutions.

**STEP 2** Rearrange the dilution expression to solve for the unknown quantity.

**STEP 3** Substitute the known quantities into the dilution expression and calculate.

---

**STEP 2** Rearrange the dilution expression to solve for the unknown quantity.

\[
\frac{C_1 V_1}{V_2} = \frac{C_2 V_2}{C_1}
\]

Divide both sides by \(C_1\)

\[
V_1 = V_2 \times \frac{C_2}{C_1}
\]

**STEP 3** Substitute the known quantities into the dilution expression and calculate.

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>(C_1 = 50.0%\ (m/v))</td>
<td>(C_2 = 35.0%\ (m/v))</td>
<td>(C_1 V_1 = C_2 V_2)</td>
</tr>
<tr>
<td>(V_1 = 1000.\ mL)</td>
<td>(V_2 = 1000.\ mL)</td>
<td>(V_1 ) decreases</td>
</tr>
</tbody>
</table>

When the final volume \(V_2\) is multiplied by a ratio of the percent concentrations (concentration factor) that is less than 1, the initial volume \(V_1\) is less than the final volume \(V_2\) as predicted in Step 1.

**STUDY CHECK 12.12**

What initial volume of a 15\% (m/v) mannose solution is needed to prepare 125 mL of a 3.0\% (m/v) mannose solution?

**ANSWER**

25 mL of a 15\% (m/v) mannose solution

**SAMPLE PROBLEM 12.13 Molarity of a Diluted Solution**

What is the molarity of a solution when 75.0 mL of a 4.00 M KCl solution is diluted to a volume of 500. mL?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** Prepare a table of the concentrations and volumes of the solutions.

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>(C_1 = 4.00\ M)</td>
<td>(C_2)</td>
<td>(C_1 V_1 = C_2 V_2)</td>
</tr>
<tr>
<td>(V_1 = 75.0\ mL)</td>
<td>(V_2 = 500.\ mL)</td>
<td>(V_2) increases</td>
</tr>
<tr>
<td>(= 0.0750\ L)</td>
<td>(= 0.500\ L)</td>
<td>(C_2) decreases</td>
</tr>
</tbody>
</table>

**STEP 2** Rearrange the dilution expression to solve for the unknown quantity.

\[
\frac{C_1 V_1}{V_2} = \frac{C_2 V_2}{V_2}
\]

Divide both sides by \(V_2\)

\[
C_2 = C_1 \times \frac{V_1}{V_2}
\]
STEP 3 Substitute the known quantities into the dilution expression and calculate.

\[ C_2 = 4.00 \text{ M} \times \frac{75.0 \text{ mL}}{500. \text{ mL}} = 0.600 \text{ M (diluted KCl solution)} \]

Volume factor decreases concentration

When the initial molarity \((C_1)\) is multiplied by a ratio of the volumes (volume factor) that is less than 1, the molarity of the diluted solution decreases as predicted in Step 1.

STUDY CHECK 12.13
What is the molarity of a solution when 50.0 mL of a 4.00 M KOH solution is diluted to 200. mL?

ANSWER
1.00 M KOH solution

12.51 To make tomato soup, you add one can of water to the condensed soup. Why is this a dilution?

12.52 A can of frozen lemonade calls for the addition of three cans of water to make a pitcher of the beverage. Why is this a dilution?

12.53 Calculate the final concentration of each of the following:
   a. 2.0 L of a 6.0 M HCl solution is added to water so that the final volume is 6.0 L.
   b. Water is added to 0.50 L of a 12 M NaOH solution to make 3.0 L of a diluted NaOH solution.
   c. A 10.0-mL sample of a 25% (m/v) KOH solution is diluted with water so that the final volume is 100.0 mL.
   d. A 50.0-mL sample of a 15% (m/v) H₂SO₄ solution is added to water to give a final volume of 250 mL.

12.54 Calculate the final concentration of each of the following:
   a. 1.0 L of a 4.0 M HNO₃ solution is added to water so that the final volume is 8.0 L.
   b. Water is added to 0.25 L of a 6.0 M NaF solution to make 2.0 L of a diluted NaF solution.
   c. A 50.0-mL sample of an 8.0% (m/v) KBr solution is diluted with water so that the final volume is 200.0 mL.
   d. A 5.0-mL sample of a 50.0% (m/v) acetic acid (CH₃CO₂H) solution is added to water to give a final volume of 25 mL.

12.55 Determine the final volume, in milliliters, of each of the following:
   a. a 1.5 M HCl solution prepared from 20.0 mL of a 6.0 M HCl solution
   b. a 2.0% (m/v) LiCl solution prepared from 50.0 mL of a 10.0% (m/v) LiCl solution
   c. a 0.500 M H₃PO₄ solution prepared from 50.0 mL of a 6.00 M H₃PO₄ solution
   d. a 5.0% (m/v) glucose solution prepared from 75 mL of a 12% (m/v) glucose solution

12.56 Determine the final volume, in milliliters, of each of the following:
   a. a 1.00% (m/v) H₂SO₄ solution prepared from 10.0 mL of a 20.0% H₂SO₄ solution
   b. a 0.10 M HCl solution prepared from 25 mL of a 6.0 M HCl solution
   c. a 1.0 M NaOH solution prepared from 50.0 mL of a 12 M NaOH solution
   d. a 1.0% (m/v) CaCl₂ solution prepared from 18 mL of a 4.0% (m/v) CaCl₂ solution

12.57 Determine the initial volume, in milliliters, required to prepare each of the following:
   a. 255 mL of a 0.200 M HNO₃ solution using a 4.00 M HNO₃ solution
   b. 715 mL of a 0.100 M MgCl₂ solution using a 6.00 M MgCl₂ solution
   c. 0.100 L of a 0.150 M KCl solution using an 8.00 M KCl solution

12.58 Determine the initial volume, in milliliters, required to prepare each of the following:
   a. 20.0 mL of a 0.250 M KNO₃ solution using a 6.00 M KNO₃ solution
   b. 25.0 mL of a 2.50 M H₂SO₄ solution using a 12.0 M H₂SO₄ solution
   c. 0.500 L of a 1.50 M NH₄Cl solution using a 10.0 M NH₄Cl solution

Applications

12.59 You need 500. mL of a 5.0% (m/v) glucose solution. If you have a 25% (m/v) glucose solution on hand, how many milliliters do you need?

12.60 A doctor orders 100. mL of 2.0% (m/v) ibuprofen. If you have 8.0% (m/v) ibuprofen on hand, how many milliliters do you need?
12.6 Chemical Reactions in Solution

**LEARNING GOAL** Given the volume and concentration of a solution in a chemical reaction, calculate the amount of a reactant or product in the reaction.

When chemical reactions involve aqueous solutions, we use the balanced chemical equation, the molarity, and the volume to determine the moles or grams of a reactant or product. For example, we can determine the volume of a solution from the molarity and the grams of reactant as seen in Sample Problem 12.14.

**SAMPLE PROBLEM 12.14 Volume of a Solution in a Reaction**
Zinc reacts with HCl to produce hydrogen gas, H₂, and ZnCl₂.

$$\text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{H}_2(g) + \text{ZnCl}_2(aq)$$

How many liters of a 1.50 M HCl solution completely react with 5.32 g of zinc?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>5.32 g Zn, 1.50 M HCl solution</td>
<td>molar mass of Zn, molarity, molar–mole factor</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2 mol HCl</td>
<td>liters of HCl solution</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to calculate the needed quantity.

grams of Zn \( \rightarrow \) Molar mass \( \rightarrow \) moles of Zn \( \rightarrow \) Mole–mole factor \( \rightarrow \) moles of HCl

Molarity \( \rightarrow \) liters of HCl solution

**STEP 3** Write equalities and conversion factors including mole–mole and concentration factors.

$$\frac{1 \text{ mol of Zn}}{65.41 \text{ g Zn}} \times \frac{65.41 \text{ g Zn}}{1 \text{ mol Zn}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Zn}} \times \frac{1 \text{ L solution}}{1.50 \text{ mol HCl}} = 0.108 \text{ L of HCl solution}$$

**STEP 4** Set up the problem to calculate the needed quantity.

$$5.32 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.41 \text{ g Zn}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Zn}} \times \frac{1 \text{ L solution}}{1.50 \text{ mol HCl}} = 0.108 \text{ L of HCl solution}$$

**STUDY CHECK 12.14**
Using the reaction in Sample Problem 12.14, how many grams of zinc can react with 225 mL of a 0.200 M HCl solution?

**ANSWER**
1.47 g of Zn
SAMPLE PROBLEM 12.15 Volume of a Reactant in a Solution

How many milliliters of a 0.250 M BaCl₂ solution are needed to react with 0.0325 L of a 0.160 M Na₂SO₄ solution?

Na₂SO₄(aq) + BaCl₂(aq) → BaSO₄(s) + 2NaCl(aq)

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Needed</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0325 L of 0.160 M Na₂SO₄ solution,</td>
<td>milliliters of BaCl₂</td>
<td>mole–mole factor,</td>
</tr>
<tr>
<td>0.250 M BaCl₂ solution</td>
<td>metric factor</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Write a plan to calculate the needed quantity.

Molarity

STEP 3 Write equalities and conversion factors including mole–mole and concentration factors.

\[
\begin{align*}
1 \text{ L of solution} &= 0.160 \text{ mol of Na₂SO₄} \\
0.160 \text{ mol Na₂SO₄} &\quad \text{and} \quad 1 \text{ L solution} = 0.160 \text{ mol Na₂SO₄} \\
0.250 \text{ mol BaCl₂} &\quad \text{and} \quad 1 \text{ L solution} = 0.250 \text{ mol BaCl₂} \\
1 \text{ mol of Na₂SO₄} &= 1 \text{ mol of BaCl₂} \\
1 \text{ mol BaCl₂} &\quad \text{and} \quad 1 \text{ mol Na₂SO₄} \\
1000 \text{ mL} &\quad \text{and} \quad 1 \text{ L} \\
&\quad \text{and} \quad 1 \text{ L} \text{ of solution} = 1000 \text{ mL} \\
&\quad \text{and} \quad 1 \text{ L} \\
&\quad \text{and} \quad 1000 \text{ mL} \\
\end{align*}
\]

STEP 4 Set up the problem to calculate the needed quantity.

\[
0.0325 \text{ L solution} \times \frac{0.160 \text{ mol Na₂SO₄}}{1 \text{ L solution}} \times \frac{1 \text{ mol BaCl₂}}{0.250 \text{ mol BaCl₂}} \times \frac{1 \text{ L solution}}{0.160 \text{ mol Na₂SO₄}} \times \frac{1000 \text{ mL BaCl₂ solution}}{1 \text{ L solution}} = 20.8 \text{ mL of BaCl₂ solution}
\]

STUDY CHECK 12.15

For the reaction in Sample Problem 12.15, how many milliliters of a 0.330 M Na₂SO₄ solution are needed to react with 26.8 mL of a 0.216 M BaCl₂ solution?

ANSWER

17.5 mL of Na₂SO₄ solution
SAMPLE PROBLEM 12.16 Volume of a Gas from a Reaction in Solution

Acid rain results from the reaction of nitrogen dioxide with water in the air.

\[ 3\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq) + \text{NO}(g) \]

At STP, how many liters of NO\(_2\) gas are required to produce 0.275 L of a 0.400 M HNO\(_3\) solution?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Equation</strong></td>
<td>0.275 L of 0.400 M HNO(_3) solution</td>
<td>liters of NO(_2)(g) at STP</td>
<td>mole–mole factor, molar volume</td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to calculate the needed quantity. We start the problem with the volume and molarity of the HNO\(_3\) solution to calculate moles. Then we can use the mole–mole factor and the molar volume to calculate the liters of NO\(_2\) gas.

**STEP 3** Write equalities and conversion factors including mole–mole and concentration factors.

\[
\begin{align*}
1 \text{ L of solution} & = 0.400 \text{ mol of HNO}_3 \\
0.400 \text{ mol HNO}_3 & = 1 \text{ L solution} \\
3 \text{ mol of NO}_2 & = 2 \text{ mol of HNO}_3 \\
3 \text{ mol NO}_2 & = 2 \text{ mol HNO}_3 \\
2 \text{ mol HNO}_3 & = 3 \text{ mol NO}_2 \\
1 \text{ mol NO}_2 & = 22.4 \text{ L NO}_2 \text{ (STP)} \\
22.4 \text{ L NO}_2 \text{ (STP)} & = 1 \text{ mol NO}_2 \\
22.4 \text{ L NO}_2 \text{ (STP)} & = 1 \text{ mol NO}_2 \\
0.275 \text{ L solution} & = 0.400 \text{ mol HNO}_3
\end{align*}
\]

**STEP 4** Set up the problem to calculate the needed quantity.

\[
0.275 \text{ L solution} \times \frac{0.400 \text{ mol HNO}_3}{1 \text{ L solution}} \times \frac{3 \text{ mol NO}_2}{2 \text{ mol HNO}_3} \times \frac{22.4 \text{ L NO}_2 \text{ (STP)}}{1 \text{ mol NO}_2} = 3.70 \text{ L of NO}_2 \text{ (STP)}
\]

**STUDY CHECK 12.16**

Using the equation in Sample Problem 12.16, determine the volume of NO produced at 100 °C and 1.20 atm, when 2.20 L of a 1.50 M HNO\(_3\) solution is produced.

**ANSWER**

42.1 L of NO
FIGURE 12.9 gives a summary of the pathways and conversion factors needed for substances including solutions involved in chemical reactions.

![Diagram](image)

FIGURE 12.9 ▶ In calculations involving chemical reactions, substance A is converted to moles of A using molar mass (if solid), gas laws (if gas), or molarity (if solution). Then moles of A are converted to moles of substance B, which are converted to grams of solid, liters of gas, or liters of solution, as needed.

What sequence of conversion factors would you use to calculate the number of grams of CaCO₃ needed to react with 1.50 L of a 2.00 M HCl solution in the reaction CaCO₃(s) + 2HCl(aq)?

### QUESTIONS AND PROBLEMS

#### 12.6 Chemical Reactions in Solution

**LEARNING GOAL** Given the volume and concentration of a solution in a chemical reaction, calculate the amount of a reactant or product in the reaction.

12.64 Answer the following for the reaction:

\[ \text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{CaCl}_2(aq) \]

a. How many milliliters of a 0.200 M HCl solution can react with 8.25 g of CaCO₃?

b. How many liters of CO₂ gas can form at STP when 15.5 mL of a 0.300 M HCl solution reacts with excess CaCO₃?

c. What is the molarity of a HCl solution if the reaction of 200. mL of the HCl solution with excess CaCO₃ produces 12.0 L of CO₂ gas at 725 mmHg and 18 °C?

12.65 Answer the following for the reaction:

\[ \text{Zn}(s) + 2\text{HBr}(aq) \rightarrow \text{H}_2(g) + \text{ZnBr}_2(aq) \]

a. How many milliliters of a 3.50 M HBr solution are required to react with 8.56 g of zinc?

b. How many liters of hydrogen gas can form at STP when 0.750 L of a 6.00 M HBr solution reacts with excess zinc?

c. What is the molarity of a HBr solution if the reaction of 28.7 mL of the HBr solution with excess zinc produces 0.620 L of H₂ gas at 725 mmHg and 24 °C?

12.66 Answer the following for the reaction:

\[ 3\text{AgNO}_3(aq) + \text{Na}_3\text{PO}_4(aq) \rightarrow \text{Ag}_3\text{PO}_4(s) + 3\text{NaNO}_3(aq) \]

a. How many milliliters of a 0.265 M AgNO₃ solution are required to react with 31.2 mL of 0.154 M Na₃PO₄ solution?

b. How many grams of silver phosphate are produced from the reaction of 25.0 mL of a 0.165 M AgNO₃ solution and excess Na₃PO₄?

c. What is the molarity of 20.0 mL of a Na₃PO₄ solution that reacts completely with 15.0 mL of a 0.576 M AgNO₃ solution?
The size and number of solute particles in different types of mixtures play an important role in determining the properties of that mixture.

Solutions
In the solutions discussed up to now, the solute was dissolved as small particles that are uniformly dispersed throughout the solvent to give a homogeneous solution. When you observe a solution, such as salt water, you cannot visually distinguish the solute from the solvent. The solution appears transparent, although it may have a color. The particles are so small that they go through filters and through semipermeable membranes. A semipermeable membrane allows solvent molecules such as water and very small solute particles to pass through, but does allow the passage of large solute molecules.

Colloids
The particles in a colloid are much larger than solute particles in a solution. Colloidal particles are large molecules, such as proteins, or groups of molecules or ions. Colloids, similar to solutions, are homogeneous mixtures that do not separate or settle out. Colloidal particles are small enough to pass through filters, but too large to pass through semipermeable membranes. Table 12.9 lists several examples of colloids.

<table>
<thead>
<tr>
<th>Table 12.9: Examples of Colloids</th>
</tr>
</thead>
<tbody>
<tr>
<td>Colloid</td>
</tr>
<tr>
<td>Fog, clouds, hair sprays</td>
</tr>
<tr>
<td>Dust, smoke</td>
</tr>
<tr>
<td>Shaving cream, whipped cream, soapsuds</td>
</tr>
<tr>
<td>Styrofoam, marshmallows</td>
</tr>
<tr>
<td>Mayonnaise, homogenized milk</td>
</tr>
<tr>
<td>Cheese, butter</td>
</tr>
<tr>
<td>Blood plasma, paints (latex), gelatin</td>
</tr>
</tbody>
</table>

Suspensions
Suspensions are heterogeneous, nonuniform mixtures that are very different from solutions or colloids. The particles of a suspension are so large that they can often be seen with the naked eye. They are trapped by filters and semipermeable membranes.

The weight of the suspended solute particles causes them to settle out soon after mixing. If you stir muddy water, it mixes but then quickly separates as the suspended particles settle to the bottom and leave clear liquid at the top. You can find suspensions among the medications in a hospital or in your medicine cabinet. These include Kaopectate, calamine lotion, antacid mixtures, and liquid penicillin. It is important to follow the instructions on the label that states “shake well before using” so that the particles form a suspension.

Water-treatment plants make use of the properties of suspensions to purify water. When chemicals such as aluminum sulfate or iron(III) sulfate are added to untreated water, they react with impurities to form large suspended particles called floc. In the water-treatment plant, a system of filters traps the suspended particles, but clean water passes through.

Table 12.10 compares the different types of mixtures and Figure 12.10 illustrates some properties of solutions, colloids, and suspensions.
TABLE 12.10 Comparison of Solutions, Colloids, and Suspensions

<table>
<thead>
<tr>
<th>Type of Mixture</th>
<th>Type of Particle</th>
<th>Settling</th>
<th>Separation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solution</td>
<td>Small particles such as atoms, ions, or small molecules</td>
<td>Particles do not settle</td>
<td>Particles cannot be separated by filters or semipermeable membranes</td>
</tr>
<tr>
<td>Colloid</td>
<td>Larger molecules or groups of molecules or ions</td>
<td>Particles do not settle</td>
<td>Particles can be separated by semipermeable membranes but not by filters</td>
</tr>
<tr>
<td>Suspension</td>
<td>Very large particles that may be visible</td>
<td>Particles settle rapidly</td>
<td>Particles can be separated by filters</td>
</tr>
</tbody>
</table>

A filter can be used to separate suspension particles from a solution, but a semipermeable membrane is needed to separate colloids from a solution. Explain.

FIGURE 12.10 Properties of different types of mixtures: (a) suspensions settle out; (b) suspensions are separated by a filter; (c) solution particles go through a semipermeable membrane, but colloids and suspensions do not.

Freezing Point Lowering and Boiling Point Elevation

When we add a solute to water, it changes the vapor pressure, freezing point, and boiling point of the solvent pure water. These types of changes in physical properties, known as colligative properties, depend only on the concentration of solute particles in the solution.

We can illustrate how these changes in physical properties occur by comparing the number of evaporating solvent molecules in pure solvent with those in a solution with a nonvolatile solute in the solvent. In the solution, there are fewer solvent molecules at the surface because the solute that has been added takes up some of the space at the surface. As a result, fewer solvent molecules can evaporate compared to the pure solvent. Vapor pressure is lowered for the solution. If we add more solute molecules, the vapor pressure will be lowered even more.

The freezing point of a solvent is lowered when a nonvolatile solute is added. In this case, the solute particles prevent the organization of solvent molecules needed to form the solid state. Thus, a lower temperature is required before the molecules of the solvent can become organized enough to freeze.

Insects and fish in climates with subfreezing temperatures control ice formation by producing biological antifreezes made of glycerol, proteins, and sugars, such as glucose, within their bodies. Some insects can survive temperatures below −60 °C. These forms of biological antifreezes may one day be applied to the long-term preservation of human organs.

The boiling point of a solvent is raised when a nonvolatile solute is added. The vapor pressure of a solvent must reach atmospheric pressure before it begins boiling. However, because the solute lowers the vapor pressure of the solvent, a temperature higher than the normal boiling point of the pure solvent is needed to cause the solution to boil. When you spread salt on an icy sidewalk when temperatures drop below freezing, the particles from

The Alaska Upis beetle produces biological antifreeze to survive subfreezing temperatures.
the salt combine with water to lower the freezing point, which causes the ice to melt. Another example is the addition of antifreeze, such as ethylene glycol, \( \text{C}_2\text{H}_6\text{O}_2 \), to the water in a car radiator. If an ethylene glycol and water mixture is about 50–50% by mass, it does not freeze until the temperature drops to about \(-30\, ^\circ\text{F}\), and does not boil unless the temperature reaches about \(225\, ^\circ\text{F}\). The solution in the radiator prevents the water in the radiator from forming ice in cold weather or boiling over on a hot desert highway.

**Particles in Solution**

A solute that is a nonelectrolyte dissolves as molecules, whereas a solute that is a strong electrolyte dissolves entirely as ions. A common solute in antifreeze is the nonelectrolyte ethylene glycol, \( \text{C}_2\text{H}_6\text{O}_2 \).

Nonelectrolyte: 1 mol of \( \text{C}_2\text{H}_6\text{O}_2(l) \) = 1 mol of \( \text{C}_2\text{H}_6\text{O}_2(aq) \)

However, when 1 mol of a strong electrolyte, such as NaCl or CaCl\(_2\), dissolves in water, the NaCl solution will contain 2 mol of particles and the CaCl\(_2\) solution will contain 3 mol of particles.

Strong electrolytes:

\[
1 \text{ mol of NaCl} = 1 \text{ mol of Na}^+(aq) + 1 \text{ mol of Cl}^-(aq)
\]

\[
2 \text{ mol of particles (aq)}
\]

\[
1 \text{ mol of CaCl}_2 = 1 \text{ mol of Ca}^{2+}(aq) + 2 \text{ mol of Cl}^-(aq)
\]

\[
3 \text{ mol of particles (aq)}
\]

**Molality (m)**

The calculation for freezing point lowering or boiling point elevation uses a concentration unit of molality. The molality, abbreviation \( m \), of a solution is the number of moles of solute particles per kilogram of solvent. This may seem similar to molarity, but the denominator for molality refers to the mass of the solvent, not the volume of the solution.

\[
\text{Molality (m)} = \frac{\text{moles of solute particles}}{\text{kilograms of solvent}}
\]

**SAMPLE PROBLEM 12.17 Calculating Molality**

Calculate the molality of a solution containing 35.5 g of the nonelectrolyte glucose \( (\text{C}_6\text{H}_{12}\text{O}_6) \) in 0.400 kg of water.

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>35.5 g of glucose (( \text{C}<em>6\text{H}</em>{12}\text{O}_6 )), 0.400 kg of water</td>
<td>molality (( m ))</td>
<td>molar mass of glucose, ( \frac{\text{moles}}{\text{kg}} )</td>
</tr>
</tbody>
</table>

**STEP 2** Write the molality expression.

\[
\text{Molality (m) of glucose solution} = \frac{\text{moles of glucose}}{\text{kilograms of water}}
\]
**Core Chemistry Skill**
Calculating the Freezing Point/Boiling Point of a Solution

**Engage**
Why does 1 mol solution of KBr, a strong electrolyte, lower the freezing point more than 1 mol solution of urea, a nonelectrolyte?

**Sample Problem 12.18 Calculating Freezing Point of a Solution**

In the northeastern United States during freezing weather, CaCl₂ is spread on icy highways to melt the ice. Calculate the freezing point lowering and freezing point of a solution containing 225 g of CaCl₂ in 500. g of water.

**Try It First**

**Solution**

**Step 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>225 g of CaCl₂,</td>
<td>ΔTᵥ, freezing point</td>
<td>molar mass, $ΔTᵥ = m × Kᵢ$</td>
<td></td>
</tr>
<tr>
<td>500. g of water</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>= 0.500 kg of water</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
STEP 2  Determine the number of moles of solute particles and calculate the molality.

We use molar mass to calculate the moles of \( \text{CaCl}_2 \). Then we multiply by three to obtain the number of moles of ions (particles) produced by 1 mol of \( \text{CaCl}_2 \) in solution.

\[
\text{moles of particles} = \frac{225 \text{ g} \cdot \text{CaCl}_2}{110.98 \text{ g} \cdot \text{CaCl}_2} \times \frac{3 \text{ mol particles}}{1 \text{ mol} \cdot \text{CaCl}_2} = 6.08 \text{ mol of particles}
\]

The molality \( (m) \) of the particles in solution is obtained by dividing the moles of particles by the number of kilograms of water in the solution.

\[
m = \frac{\text{moles of particles}}{\text{kilograms of water}} = \frac{6.08 \text{ mol particles}}{0.500 \text{ kg water}} = 12.2 \text{ m}
\]

STEP 3  Calculate the temperature change and subtract from the freezing point. The freezing point lowering is calculated using the molality and the freezing point constant. Finally, the freezing point lowering is subtracted from 0.00 °C to obtain the new freezing point of the \( \text{CaCl}_2 \) solution.

\[
\Delta T_f = m \times K_f
\]

\[
\Delta T_f = 12.2 \text{ m} \times \frac{1.86 \degree \text{C}}{\text{m}} = 22.7 \degree \text{C}
\]

\[
T_{\text{solution}} = T_{\text{water}} - \Delta T_f
\]

\[
T_{\text{solution}} = 0.0 \degree \text{C} - 22.7 \degree \text{C} = -22.7 \degree \text{C}
\]

STUDY CHECK 12.18

Ethylene glycol, \( \text{C}_2\text{H}_6\text{O}_2 \), a nonelectrolyte, is added to the water in a radiator to give a solution containing 515 g of ethylene glycol in 565 g of water (solvent). Calculate the freezing point lowering and freezing point of the solution.

ANSWER

\( \Delta T_f = 27.2 \degree \text{C} \); freezing point is \(-27.2 \degree \text{C}\)

Boiling Point Elevation

A similar change occurs with the boiling point of water. The boiling point elevation \( (\Delta T_b) \) is determined from the molality \( (m) \) of the particles in the solution and the boiling point constant, \( K_b \), which is \( \frac{0.52 \degree \text{C}}{m} \) for water.

\[
\Delta T_b = m \times K_b
\]

SAMPLE PROBLEM 12.19  Calculating the Boiling Point of a Solution

Propylene glycol, \( \text{C}_3\text{H}_8\text{O}_2 \), is a nonelectrolyte that is added to the water in a radiator to increase the boiling point. If 4.6 mol of propylene glycol is added to 1.55 kg of water (solvent) in a radiator, what is the boiling point, in degrees Celsius, of the solution?
**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>4.6 mol of propylene glycol, 1.55 kg of water</td>
<td>boiling point</td>
<td>$\Delta T_b = m \times K_b$</td>
</tr>
</tbody>
</table>

**STEP 2** Determine the number of moles of solute particles and calculate the molality. Since propylene glycol is a nonelectrolyte, the moles of propylene glycol is equal to the moles of particles.

$$m = \frac{\text{moles of particles}}{\text{kilograms of water}} = \frac{4.6 \text{ mol}}{1.55 \text{ kg}} = 3.0 \text{ m}$$

**STEP 3** Calculate the temperature change and add to the boiling point.

$$\Delta T_b = m \times K_b = 3.0 \text{ m} \times 0.52 \, ^\circ \text{C} = 1.6 \, ^\circ \text{C}$$

$$T_{\text{solution}} = T_{\text{water}} + \Delta T_b$$

$$= 100.0 \, ^\circ \text{C} + 1.6 \, ^\circ \text{C}$$

$$= 101.6 \, ^\circ \text{C}$$

**STUDY CHECK 12.19**

A 1.2-mol sample of potassium phosphate, $\text{K}_3\text{PO}_4$, a strong electrolyte, is added to 0.26 kg of water. What is the boiling point, in degrees Celsius, of this solution?

**ANSWER**

109.6 °C

The effect of some solutions on freezing and boiling point is summarized in **TABLE 12.11**.

<p>| TABLE 12.11 Effect of Solute Concentration on Freezing and Boiling Points of 1 kg of Water |
|-----------------------------------------------|-----------------------------------------------|-----------------------------------------------|-----------------------------------------------|</p>
<table>
<thead>
<tr>
<th><strong>Substance/kg water</strong></th>
<th><strong>Type of Solute</strong></th>
<th><strong>Molality of Particles</strong></th>
<th><strong>Freezing Point</strong></th>
<th><strong>Boiling Point</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>Pure water</td>
<td>None</td>
<td>0</td>
<td>0 °C</td>
<td>100 °C</td>
</tr>
<tr>
<td>1 mol of $\text{C}_2\text{H}_4\text{O}_2$</td>
<td>Nonelectrolyte</td>
<td>1 m</td>
<td>−1.86 °C</td>
<td>100.52 °C</td>
</tr>
<tr>
<td>1 mol of $\text{NaCl}$</td>
<td>Strong electrolyte</td>
<td>2 m</td>
<td>−3.72 °C</td>
<td>101.04 °C</td>
</tr>
<tr>
<td>1 mol of $\text{CaCl}_2$</td>
<td>Strong electrolyte</td>
<td>3 m</td>
<td>−5.58 °C</td>
<td>101.56 °C</td>
</tr>
</tbody>
</table>

**QUESTIONS AND PROBLEMS**

**12.7 Molality and Freezing Point Lowering/Boiling Point Elevation**

**LEARNING GOAL** Identify a mixture as a solution, a colloid, or a suspension. Using the molality, calculate the new freezing point and new boiling point for a solution.

**12.67** Identify the following as characteristic of a solution, colloid, or suspension:

- a. a mixture that cannot be separated by a semipermeable membrane
- b. a mixture that settles out upon standing

**12.68** Identify the following as characteristic of a solution, colloid, or suspension:

- a. Particles of this mixture remain inside a semipermeable membrane but pass through filters.
- b. The particles of solute in this mixture are very large and visible.
12.69 In each pair, identify the solution that will have a lower freezing point. Explain.
   a. 1.0 mol of glycerol (nonelectrolyte) and 2.0 mol of ethylene glycol (nonelectrolyte) each in 1.0 kg of water
   b. 0.50 mol of KCl (strong electrolyte) and 0.50 mol of MgCl₂ (strong electrolyte) each in 1.0 kg of water

12.70 In each pair, identify the solution that will have a higher boiling point. Explain.
   a. 1.50 mol of LiOH (strong electrolyte) and 3.00 mol of KOH (strong electrolyte) each in 1.0 kg of water
   b. 0.40 mol of Al(NO₃)₃ (strong electrolyte) and 0.40 mol of CsCl (strong electrolyte) each in 1.0 kg of water

12.71 Calculate the molality (m) of the following solutions:
   a. 325 g of methanol, CH₃OH, a nonelectrolyte, added to 455 g of water
   b. 640. g of the antifreeze propylene glycol, C₃H₈O₂, a nonelectrolyte, dissolved in 1.22 kg of water

12.72 Calculate the molality (m) of the following solutions:
   a. 65.0 g of glucose, C₆H₁₂O₆, a nonelectrolyte, dissolved in 112 g of water
   b. 110. g of sucrose, C₁₂H₂₂O₁₁, a nonelectrolyte, dissolved in 1.50 kg of water

12.73 Calculate the freezing point and boiling point of each solution in problem 12.71.

12.74 Calculate the freezing point and boiling point of each solution in problem 12.72.

12.8 Properties of Solutions: Osmosis

**LEARNING GOAL** Describe how the number of particles in solution affects osmosis.

The movement of water into and out of the cells of plants as well as the cells of our bodies is an important biological process that also depends on the solute concentration. In a process called **osmosis**, water molecules move through a semipermeable membrane from the solution with the lower concentration of solute into a solution with the higher solute concentration. In an osmosis apparatus, water is placed on one side of a semipermeable membrane and a sucrose (sugar) solution on the other side. The semipermeable membrane allows water molecules to flow back and forth but blocks the sucrose molecules because they cannot pass through the membrane. Because the sucrose solution has a higher solute concentration, more water molecules flow into the sucrose solution than out of the sucrose solution. The volume level of the sucrose solution rises as the volume level on the water side falls. The increase of water dilutes the sucrose solution to equalize (or attempt to equalize) the concentrations on both sides of the membrane.

Eventually the height of the sucrose solution creates sufficient pressure to equalize the flow of water between the two compartments. This pressure, called **osmotic pressure**, prevents the flow of additional water into the more concentrated solution. Then there is no

![Semipermeable membrane diagram](image-url)

Water flows into the solution with a higher solute concentration until the flow of water becomes equal in both directions.
further change in the volumes of the two solutions. The osmotic pressure depends on the concentration of solute particles in the solution. The greater the number of particles dissolved, the higher its osmotic pressure. In this example, the sucrose solution has a higher osmotic pressure than pure water, which has an osmotic pressure of zero.

In a process called reverse osmosis, a pressure greater than the osmotic pressure is applied to a solution so that it is forced through a purification membrane. The flow of water is reversed because water flows from an area of lower water concentration to an area of higher water concentration. The molecules and ions in solution stay behind, trapped by the membrane, while water passes through the membrane. This process of reverse osmosis is used in a few desalination plants to obtain pure water from sea (salt) water. However, the pressure that must be applied requires so much energy that reverse osmosis is not yet an economical method for obtaining pure water in most parts of the world.

Isotonic Solutions

Because the cell membranes in biological systems are semipermeable, osmosis is an ongoing process. The solutes in body solutions such as blood, tissue fluids, lymph, and plasma all exert osmotic pressure. Most intravenous (IV) solutions used in a hospital are isotonic solutions, which exert the same osmotic pressure as body fluids such as blood. The percent concentration typically used in IV solutions is mass/volume percent \((m/v)\), which is a type of percent concentration we have already discussed. The most typical isotonic solutions are 0.9\% \((m/v)\) NaCl solution, or 0.9 g of NaCl/100. mL of solution, and 5\% \((m/v)\) glucose, or 5 g of glucose/100. mL of solution. Although they do not contain the same kinds of particles, a 0.9\% \((m/v)\) NaCl solution as well as a 5\% \((m/v)\) glucose solution both have the same osmotic pressure. A red blood cell placed in an isotonic solution retains its volume because there is an equal flow of water into and out of the cell (see FIGURE 12.11a).

Hypotonic and Hypertonic Solutions

If a red blood cell is placed in a solution that is not isotonic, the differences in osmotic pressure inside and outside the cell can drastically alter the volume of the cell. When a red blood cell is placed in a hypotonic solution, which has a lower solute concentration (hypo means “lower than”), water flows into the cell by osmosis. The increase in fluid causes the cell to swell, and possibly burst—a process called hemolysis (see FIGURE 12.11b). A similar process occurs when you place dehydrated food, such as raisins or dried fruit, in water. The water enters the cells, and the food becomes plump and smooth.

If a red blood cell is placed in a hypertonic solution, which has a higher solute concentration (hyper means “greater than”), water flows out of the cell into the hypertonic

---

**FIGURE 12.11** (a) In an isotonic solution, a red blood cell retains its normal volume. (b) Hemolysis: In a hypotonic solution, water flows into a red blood cell, causing it to swell and burst. (c) Crenation: In a hypertonic solution, water leaves the red blood cell, causing it to shrink.

**What happens to a red blood cell placed in a 4% NaCl solution?**
solution by osmosis. Suppose a red blood cell is placed in a 10% (m/v) NaCl solution. Because the osmotic pressure in the red blood cell is the same as a 0.9% (m/v) NaCl solution, the 10% (m/v) NaCl solution has a much greater osmotic pressure. As water leaves the cell, it shrinks, a process called crenation (see FIGURE 12.11c). A similar process occurs when making pickles, which uses a hypertonic salt solution that causes the cucumbers to shrink as they lose water.

SAMPLE PROBLEM 12.20 Isotonic, Hypotonic, and Hypertonic Solutions

Describe each of the following solutions as isotonic, hypotonic, or hypertonic. Indicate whether a red blood cell placed in each solution will undergo hemolysis, crenation, or no change.

a. a 5% (m/v) glucose solution
b. a 0.2% (m/v) NaCl solution

TRY IT FIRST

SOLUTION

a. A 5% (m/v) glucose solution is isotonic. A red blood cell will not undergo any change.
b. A 0.2% (m/v) NaCl solution is hypotonic. A red blood cell will undergo hemolysis.

STUDY CHECK 12.20

What will happen to a red blood cell placed in a 10% (m/v) glucose solution?

ANSWER

The red blood cell will shrink (crenate).

Dialysis

Dialysis is a process that is similar to osmosis. In dialysis, a semipermeable membrane, called a dialyzing membrane, permits small solute molecules and ions as well as solvent water molecules to pass through, but it retains large particles, such as colloids. Dialysis is a way to separate solution particles from colloids.

Suppose we fill a cellophane bag with a solution containing NaCl, glucose, starch, and protein and place it in pure water. Cellophane is a dialyzing membrane, and the sodium ions, chloride ions, and glucose molecules will pass through it into the surrounding water. However, large colloidal particles, like starch and protein, remain inside. Water molecules will flow into the cellophane bag. Eventually the concentrations of sodium ions, chloride ions, and glucose molecules inside and outside the dialysis bag become equal. To remove more NaCl or glucose, the cellophane bag must be placed in a fresh sample of pure water.

CHEMISTRY LINK TO HEALTH

Dialysis by the Kidneys and the Artificial Kidney

The fluids of the body undergo dialysis by the membranes of the kidneys, which remove waste materials, excess salts, and water. In an adult, each kidney contains about 2 million nephrons. At the top of each nephron, there is a network of arterial capillaries called the glomerulus.

As blood flows into the glomerulus, small particles, such as amino acids, glucose, urea, water, and certain ions, will move through the capillary membranes into the nephron. As this solution moves through the nephron, substances still of value to the body (such as amino acids, glucose, certain ions, and 99% of the water) are reabsorbed. The major waste product, urea, is excreted in the urine.

Hemodialysis

If the kidneys fail to dialyze waste products, increased levels of urea can become life-threatening in a relatively short time. A person with kidney failure must use an artificial kidney, which cleanses the blood by hemodialysis.

A typical artificial kidney machine contains a large tank filled with water containing selected electrolytes. In the center of this dialyzing bath (dialysate), there is a dialyzing coil or membrane made of cellulose tubing. As the patient’s blood flows through the dialyzing coil, the highly concentrated waste products dialyze out of the
blood. No blood is lost because the membrane is not permeable to large particles such as red blood cells.

Dialysis patients do not produce much urine. As a result, they retain large amounts of water between dialysis treatments, which produces a strain on the heart. The intake of fluids for a dialysis patient may be restricted to as little as a few teaspoons of water a day. In the dialysis procedure, the pressure of the blood is increased as it circulates through the dialyzing coil so water can be squeezed out of the blood. For some dialysis patients, 2 to 10 L of water may be removed during one treatment. Dialysis patients have from two to three treatments a week, each treatment requiring about 5 to 7 h. Some of the newer treatments require less time. For many patients, dialysis is done at home with a home dialysis unit.

During dialysis, waste products and excess water are removed from the blood.

**QUESTIONS AND PROBLEMS**

### 12.8 Properties of Solutions: Osmosis

**LEARNING GOAL** Describe how the number of particles in solution affects osmosis.

**12.75** A 10% (m/v) starch solution is separated from a 1% (m/v) starch solution by a semipermeable membrane. (Starch is a colloid.)
   a. Which compartment has the higher osmotic pressure?
   b. In which direction will water flow initially?
   c. In which compartment will the volume level rise?

**12.76** A 0.1% (m/v) albumin solution is separated from a 2% (m/v) albumin solution by a semipermeable membrane. (Albumin is a colloid.)
   a. Which compartment has the higher osmotic pressure?
   b. In which direction will water flow initially?
   c. In which compartment will the volume level rise?

**12.77** Indicate the compartment (A or B) that will increase in volume for each of the following pairs of solutions separated by a semipermeable membrane:

<table>
<thead>
<tr>
<th></th>
<th>A</th>
<th>B</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.</td>
<td>5% (m/v) sucrose</td>
<td>10% (m/v) sucrose</td>
</tr>
<tr>
<td>b.</td>
<td>8% (m/v) albumin</td>
<td>4% (m/v) albumin</td>
</tr>
<tr>
<td>c.</td>
<td>0.1% (m/v) starch</td>
<td>10% (m/v) starch</td>
</tr>
</tbody>
</table>

**Applications**

**12.79** Are the following solutions isotonic, hypotonic, or hypertonic compared with a red blood cell?
   a. distilled H₂O
   b. 1% (m/v) glucose
   c. 0.9% (m/v) NaCl
   d. 15% (m/v) glucose

**12.80** Will a red blood cell undergo crenation, hemolysis, or no change in each of the following solutions?
   a. 1% (m/v) glucose
   b. 2% (m/v) NaCl
   c. 5% (m/v) glucose
   d. 0.1% (m/v) NaCl

**12.78** Indicate the compartment (A or B) that will increase in volume for each of the following pairs of solutions separated by a semipermeable membrane:

<table>
<thead>
<tr>
<th></th>
<th>A</th>
<th>B</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.</td>
<td>20% (m/v) starch</td>
<td>10% (m/v) starch</td>
</tr>
<tr>
<td>b.</td>
<td>10% (m/v) albumin</td>
<td>2% (m/v) albumin</td>
</tr>
<tr>
<td>c.</td>
<td>0.5% (m/v) sucrose</td>
<td>5% (m/v) sucrose</td>
</tr>
</tbody>
</table>
12.81 Each of the following mixtures is placed in a dialyzing bag and immersed in distilled water. Which substances will be found outside the bag in the distilled water?

a. NaCl solution
b. starch solution (colloid) and alanine (an amino acid) solution
c. NaCl solution and starch solution (colloid)
d. urea solution

12.82 Each of the following mixtures is placed in a dialyzing bag and immersed in distilled water. Which substances will be found outside the bag in the distilled water?

a. KCl solution and glucose solution
b. albumin solution (colloid)
c. an albumin solution (colloid), KCl solution, and glucose solution
d. urea solution and NaCl solution

Follow Up

**USING DIALYSIS FOR RENAL FAILURE**

As a dialysis patient, Michelle has a 4-h dialysis treatment three times a week. When she arrives at the dialysis clinic, her weight, temperature, and blood pressure are taken and blood tests are done to determine the level of electrolytes and urea in her blood. In the dialysis center, tubes to the dialyzer are connected to the catheter she has had implanted. Blood is then pumped out of her body, through the dialyzer where it is filtered, and returned to her body. As Michelle’s blood flows through the dialyzer, electrolytes from the dialysate move into her blood, and waste products in her blood move into the dialysate, which is continually renewed. To achieve normal serum electrolyte levels, dialysate fluid contains sodium, chloride, and magnesium levels that are equal to serum concentrations. These electrolytes are removed from the blood only if their concentrations are higher than normal. Typically, in dialysis patients, the potassium ion level is higher than normal. Therefore, initial dialysis may start with a low concentration of potassium ion in the dialysate. During dialysis excess fluid is removed by osmosis. A 4-h dialysis session requires at least 120 L of dialysis fluid. During dialysis, the electrolytes in the dialysate are adjusted until the electrolytes have the same levels as normal serum. Initially the dialysate solution prepared for Michelle contains the following: $\text{HCO}_3^-$, $\text{K}^+$, $\text{Na}^+$, $\text{Ca}^{2+}$, $\text{Mg}^{2+}$, $\text{Cl}^-$, glucose.

**Applications**

12.83 A doctor orders 0.075 g of chlorpromazine, which is used to treat schizophrenia. If the stock solution is 2.5% (m/v), how many milliliters are administered to the patient?

12.84 A doctor orders 5.0 mg of compazine, which is used to treat nausea, vertigo, and migraine headaches. If the stock solution is 2.5% (m/v), how many milliliters are administered to the patient?

12.85 A CaCl$_2$ solution is given to increase blood levels of calcium. If a patient receives 5.0 mL of a 10.0% (m/v) CaCl$_2$ solution, how many grams of CaCl$_2$ were given?

12.86 An intravenous solution of mannitol is used as a diuretic to increase the loss of sodium and chloride by a patient. If a patient receives 30.0 mL of a 25% (m/v) mannitol solution, how many grams of mannitol were given?
12.1 Solutions

**LEARNING GOAL** Identify the solute and solvent in a solution; describe the formation of a solution.

- A solution forms when a solute dissolves in a solvent.
- In a solution, the particles of solute are evenly dispersed in the solvent.
- The solute and solvent may be solid, liquid, or gas.
- The polar $\text{OH}$ bond leads to hydrogen bonding between water molecules.
- An ionic solute dissolves in water, a polar solvent, because the polar water molecules attract and pull the ions into solution, where they become hydrated.
- The expression *like dissolves like* means that a polar or an ionic solute dissolves in a polar solvent while a nonpolar solute dissolves in a nonpolar solvent.

11.2 Electrolytes and Nonelectrolytes

**LEARNING GOAL** Identify solutes as electrolytes or nonelectrolytes.

- Substances that produce ions in water are called electrolytes because their solutions will conduct an electrical current.
- Strong electrolytes are completely dissociated, whereas weak electrolytes are only partially ionized.
- Nonelectrolytes are substances that dissolve in water to produce only molecules and cannot conduct electrical currents.

12.3 Solubility

**LEARNING GOAL** Define solubility; distinguish between an unsaturated and a saturated solution. Identify an ionic compound as soluble or insoluble.

- The solubility of a solute is the maximum amount of a solute that can dissolve in 100 g of solvent.
- A solution that contains the maximum amount of dissolved solute is a saturated solution.
- A solution containing less than the maximum amount of dissolved solute is unsaturated.
- An increase in temperature increases the solubility of most solids in water, but decreases the solubility of gases in water.
- Ionic compounds that are soluble in water usually contain $\text{Li}^+$, $\text{Na}^+$, $\text{K}^+$, $\text{NH}_4^+$, $\text{NO}_3^-$, or acetate, $\text{C}_2\text{H}_3\text{O}_2^-$.

12.4 Solution Concentrations

**LEARNING GOAL** Calculate the concentration of a solute in a solution; use concentration as conversion factors to calculate the amount of solute or solution.

- Mass percent expresses the mass/mass ($\text{m/m}$) ratio of the mass of solute to the mass of solution multiplied by 100%.
- Percent concentration can also be expressed as volume/volume ($\text{v/v}$) and mass/volume ($\text{m/v}$) ratios.
- Molarity is the moles of solute per liter of solution.
• In calculations of grams or milliliters of solute or solution, the concentration is used as a conversion factor.
• Molarity (mol/L) is written as conversion factors to solve for moles of solute or volume of solution.

12.5 Dilution of Solutions

LEARNING GOAL Describe the dilution of a solution; calculate the unknown concentration or volume when a solution is diluted.

• In dilution, a solvent such as water is added to a solution, which increases its volume and decreases its concentration.

12.6 Chemical Reactions in Solution

LEARNING GOAL Given the volume and concentration of a solution in a chemical reaction, calculate the amount of a reactant or product in the reaction.

• When solutions are involved in chemical reactions, the moles of a substance in solution can be determined from the volume and molarity of the solution.
• When mass, volume, and molarities of substances in a reaction are given, the balanced equation is used to determine the quantities or concentrations of other substances in the reaction.

12.7 Molality and Freezing Point Lowering/Boiling Point Elevation

LEARNING GOAL Identify a mixture as a solution, a colloid, or a suspension. Using the molality, calculate the new freezing point and new boiling point for a solution.

• Colloids contain particles that pass through most filters but do not settle out or pass through semipermeable membranes.
• Suspensions have very large particles that settle out.
• The particles in a solution lower the vapor pressure, lower the freezing point, and raise the boiling point.
• Molality is the moles of solute per kilogram of solvent, usually water.

12.8 Properties of Solutions: Osmosis

LEARNING GOAL Describe how the number of particles in solution affects osmosis.

• The particles in a solution increase the osmotic pressure.
• In osmosis, solvent (water) passes through a semipermeable membrane from a solution with a lower osmotic pressure (lower solute concentration) to a solution with a higher osmotic pressure (higher solute concentration).
• Isotonic solutions have osmotic pressures equal to that of body fluids.
• A red blood cell maintains its volume in an isotonic solution but swells in a hypotonic solution, and shrinks in a hypertonic solution.
• In dialysis, water and small solute particles pass through a dialyzing membrane, while larger particles are retained.

KEY TERMS

- colloid A mixture having particles that are moderately large. Colloids pass through filters but cannot pass through semipermeable membranes.
- concentration A measure of the amount of solute that is dissolved in a specified amount of solution.
- dialysis A process in which water and small solute particles pass through a semipermeable membrane.
- dilution A process by which water (solvent) is added to a solution to increase the volume and decrease (dilute) the solute concentration.
- electrolyte A substance that produces ions when dissolved in water; its solution conducts electricity.
- Henry’s law The solubility of a gas in a liquid is directly related to the pressure of that gas above the liquid.
- hydration The process of surrounding dissolved ions by water molecules.
- mass percent (m/m) The grams of solute in 100. g of solution.
- mass/volume percent (m/v) The grams of solute in 100. mL of solution.
- molarity (m) The number of moles of solute particles in exactly 1 kg of solvent.
- molarity (M) The number of moles of solute in exactly 1 L of solution.
- nonelectrolyte A substance that dissolves in water as molecules; its solution does not conduct an electrical current.
- osmosis The flow of a solvent, usually water, through a semipermeable membrane into a solution of higher solute concentration.
- osmotic pressure The pressure that prevents the flow of water into the more concentrated solution.
- saturated solution A solution containing the maximum amount of solute that can dissolve at a given temperature. Any additional solute will remain undissolved in the container.
- solubility The maximum amount of solute that can dissolve in 100. g of solvent, usually water, at a given temperature.
- solubility rules A set of guidelines that states whether an ionic compound is soluble or insoluble in water.
- solute The component in a solution that is present in the lesser amount.
- solution A homogeneous mixture in which the solute is made up of small particles (ions or molecules) that can pass through filters and semipermeable membranes.
- solvent The substance in which the solute dissolves; usually the component present in greater amount.
- strong electrolyte A compound that ionizes completely when it dissolves in water; its solution is a good conductor of electricity.
- suspension A mixture in which the solute particles are large enough and heavy enough to settle out and be retained by both filters and semipermeable membranes.
- unsaturated solution A solution that contains less solute than can be dissolved.
- volume percent (v/v) The milliliters of solute in 100. mL of solution.
- weak electrolyte A substance that produces only a few ions along with many molecules when it dissolves in water; its solution is a weak conductor of electricity.
Using Solubility Rules (12.3)

- Soluble ionic compounds contain Li⁺, Na⁺, K⁺, NH₄⁺, NO₃⁻, or CH₃CO₂⁻ (acetate).
- Ionic compounds containing Cl⁻, Br⁻, or I⁻ are soluble, but if they are combined with Ag⁺, Pb²⁺, or Hg₂²⁺, they are insoluble.
- Most ionic compounds containing SO₄²⁻ are soluble, but if they contain Ba²⁺, Pb²⁺, Ca²⁺, Sr²⁺, or Hg₂²⁺, they are insoluble.
- Most other ionic compounds including those containing the anions CO₃²⁻, S²⁻, PO₄³⁻, or OH⁻ are insoluble.
- To write an equation or ionic equation for the formation of an insoluble ionic compound, we write the cations and anions to identify any combination that would form an insoluble ionic compound.

Example: Determine if an ionic compound forms when solutions of CaCl₂ and K₂CO₃ are mixed. If so, write the net ionic equation for the reaction.

Answer: In the ionic equation, we show all the ions of the reactants and show the ionic compound CaCO₃ as a solid.

\[
\text{Ca}^2⁺(aq) + 2\text{Cl}^⁻(aq) + 2\text{K}^⁺(aq) + \text{CO}_3^{2⁻}(aq) \rightarrow \text{CaCO}_3(s) + 2\text{K}^⁺(aq) + 2\text{Cl}^⁻(aq)
\]

For the net ionic equation, spectator ions that appear on both sides of the equation are removed.

\[
\text{Ca}^2⁺(aq) + \text{CO}_3^{2⁻}(aq) \rightarrow \text{CaCO}_3(s)
\]

Net ionic equation

Calculating Concentration (12.4)

The amount of solute dissolved in a certain amount of solution is called the concentration of the solution.

- Mass percent (m/m) = \(\frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%\)
- Volume percent (v/v) = \(\frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%\)
- Mass/volume percent (m/v) = \(\frac{\text{grams of solute}}{\text{milliliters of solution}} \times 100\%\)
- Molarity (M) = \(\frac{\text{moles of solute}}{\text{liters of solution}}\)

Example: What is the mass/volume percent (m/v) and the molarity (M) of 225 mL (0.225 L) of a LiCl solution that contains 17.1 g of LiCl?

Answer: Mass/volume % (m/v) = \(\frac{\text{grams of solute}}{\text{milliliters of solution}} \times 100\%\)

\[
= \frac{17.1 \text{ g LiCl}}{225 \text{ mL solution}} \times 100\% = 7.60\% \ (\text{m/v}) \ \text{LiCl solution}
\]

moles of LiCl = \(\frac{17.1 \text{ g LiCl}}{42.39 \text{ g LiCl}}\) = 0.403 mol of LiCl

Molarity (M) = \(\frac{\text{moles of solute}}{\text{liters of solution}}\) = \(\frac{0.403 \text{ mol LiCl}}{0.225 \text{ L solution}}\) = 1.79 M LiCl solution

Using Concentration as a Conversion Factor (12.4)

- When we need to calculate the amount of solute or solution, we use the concentration as a conversion factor.
- For example, the concentration of a 4.50 M HCl solution means there are 4.50 mol of HCl in 1 L of HCl solution, which gives two conversion factors written as

\[
\frac{4.50 \text{ mol HCl}}{1 \text{ L solution}} \quad \text{and} \quad \frac{1 \text{ L solution}}{4.50 \text{ mol HCl}}
\]

Example: How many milliliters of a 4.50 M HCl solution will provide 41.2 g of HCl?

Answer:

\[
41.2 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{1 \text{ L solution}}{4.50 \text{ mol HCl}} \times \frac{1000 \text{ mL solution}}{1 \text{ L solution}} = 251 \text{ mL of HCl solution}
\]

Calculating the Quantity of a Reactant or Product for a Chemical Reaction in Solution (12.6)

- When chemical reactions involve aqueous solutions of reactants or products, we use the balanced chemical equation, the molarity, and the volume to determine the moles or grams of the reactants or products.

Example: How many grams of zinc metal will react with 0.315 L of a 1.20 M HCl solution?

\[
\text{Zn(s) + 2HCl(aq) \rightarrow H}_2(\text{g}) + \text{ZnCl}_2(aq)
\]

Answer:

\[
0.315 \text{ L solution} \times \frac{1.20 \text{ mol HCl}}{1 \text{ L solution}} \times \frac{1 \text{ mol Zn}}{2 \text{ mol HCl}} \times \frac{65.41 \text{ g Zn}}{1 \text{ mol Zn}} = 12.4 \text{ g of Zn}
\]

Calculating the Freezing Point/Boiling Point of a Solution (12.7)

- The particles in a solution lower the freezing point, raise the boiling point, and increase the osmotic pressure.
- The freezing point lowering (\(\Delta T_f\)) is determined from the molality (m) of the particles in the solution and the freezing point constant, \(K_f\).

\[
\Delta T_f = m \times K_f
\]

- The boiling point elevation (\(\Delta T_b\)) is determined from the molality (m) of the particles in the solution and the boiling point constant, \(K_b\).

\[
\Delta T_b = m \times K_b
\]

Example: What is the boiling point of a solution that contains 1.5 mol of the strong electrolyte KCl in 1 kg of water?

Answer: A solution of 1.5 mol of KCl in 1 kg of water, which contains 3.0 mol of particles (1.5 mol of K⁺ and 1.5 mol of Cl⁻), is a 3.0-m solution.

\[
\Delta T_b = m \times K_b = 3.0 \times \frac{0.52 ^\circ C}{\text{m}} = 1.6 ^\circ C
\]

\[
T_{\text{solution}} = T_{\text{water}} + \Delta T_b = 100.0 ^\circ C + 1.6 ^\circ C = 101.6 ^\circ C
\]
UNDERSTANDING THE CONCEPTS

The chapter sections to review are shown in parentheses at the end of each question.

12.87 Match the diagrams with the following: (12.1)
   a. a polar solute and a polar solvent
   b. a nonpolar solute and a polar solvent
   c. a nonpolar solute and a nonpolar solvent

12.88 If all the solute is dissolved in diagram 1, how would heating or cooling the solution cause each of the following changes? (12.3)
   a. 2 to 3
   b. 2 to 1

12.89 Select the diagram that represents the solution formed by a solute that is a (12.2)
   a. nonelectrolyte
   b. weak electrolyte
   c. strong electrolyte

12.90 Select the container that represents the dilution of a 5% (m/v) NaCl solution to give each of the following: (12.5)
   a. a 3% (m/v) NaCl solution
   b. a 2% (m/v) NaCl solution

12.91 Use the following ions: (12.3)
   Na⁺, Cl⁻, Ag⁺, NO₃⁻
   a. Select the beaker (1, 2, 3, or 4) that contains the products after the solutions in beakers A and B are mixed.
   b. If an insoluble ionic compound forms, write the ionic equation.
   c. If a reaction occurs, write the net ionic equation.

12.92 Use the following ions: (12.3)
   K⁺, NO₃⁻, NH₄⁺, Br⁻
   a. Select the beaker (1, 2, 3, or 4) that contains the products after the solutions in beakers A and B are mixed.
   b. If an insoluble ionic compound forms, write the ionic equation.
   c. If a reaction occurs, write the net ionic equation.

12.93 When a dried raisin is soaked in sugar-water solution, it becomes plump and smooth. Why? (12.6)
**ADDITIONAL QUESTIONS AND PROBLEMS**

12.94 What is crenation? (12.6)

12.95 A semipermeable membrane separates two compartments, A and B. If the levels in A and B are equal initially, select the diagram that illustrates the final levels in a to d: (12.8)

![Diagram](image)

<table>
<thead>
<tr>
<th>Solution in A</th>
<th>Solution in B</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. 2% (m/v) starch</td>
<td>8% (m/v) starch</td>
</tr>
<tr>
<td>b. 1% (m/v) starch</td>
<td>1% (m/v) starch</td>
</tr>
<tr>
<td>c. 5% (m/v) sucrose</td>
<td>1% (m/v) sucrose</td>
</tr>
<tr>
<td>d. 0.1% (m/v) sucrose</td>
<td>1% (m/v) sucrose</td>
</tr>
</tbody>
</table>

12.96 Select the diagram that represents the shape of a red blood cell when placed in each of the following a to e: (12.8)

![Diagram](image)

a. 0.9% (m/v) NaCl solution
b. 10% (m/v) glucose solution
c. 0.01% (m/v) NaCl solution
d. 5% (m/v) glucose solution
e. 1% (m/v) glucose solution

**12.105** If NaCl has a solubility of 36.0 g in 100. g of H₂O at 20 °C, how many grams of water are needed to prepare a saturated solution containing 80.0 g of NaCl? (12.3)

12.106 If the solid NaCl in a saturated solution of NaCl continues to dissolve, why is there no change in the concentration of the NaCl solution? (12.3)

12.107 Calculate the mass percent (m/m) of a vitamin C solution by dissolving 5.2 g of ascorbic acid (vitamin C) in 255.9 g water. (12.4)

12.108 Calculate the mass percent (m/m) of a sucrose solution by dissolving 78.4 g of sucrose in 327.5 g of water. (12.4)

12.109 How many milliliters of a 12% (v/v) propyl alcohol solution would you need to obtain 4.5 mL of propyl alcohol? (12.4)

12.110 An 80-proof brandy is a 40.0% (v/v) ethanol solution. The “proof” is twice the percent concentration of alcohol in the beverage. How many milliliters of alcohol are present in 750 mL of brandy? (12.4)

12.111 How many liters of a 12% (m/v) KOH solution would you need to obtain 86.0 g of KOH? (12.4)

12.112 How many liters of a 5.0% (m/v) glucose solution would you need to obtain 75 g of glucose? (12.4)

12.113 What is the molarity of a solution containing 5.0 g of Na₃PO₄ in 200. mL of Na₃PO₄ solution? (12.4)

12.114 What is the molarity of a solution containing 18 g of MgCl₂ in 375 mL of MgCl₂ solution? (12.4)

12.115 How many milliliters of a 1.75 M LiCl solution contain 15.2 g of LiCl? (12.4)

12.116 How many milliliters of a 1.50 M NaBr solution contain 75.0 g of NaBr? (12.4)

12.117 What volume, in mL, of 2.50 M of sodium chloride solution is needed to obtain 3.0 g of NaCl by evaporation? (12.4)

12.118 What is the concentration of an AgNO₃ solution when 1.50 g of silver nitrate is dissolved in 25.0 mL of water? (12.4)
12.119 How many grams of solute are in each of the following solutions? (12.4)
   a. 2.5 L of a 3.0 M Al(NO₃)₃ solution
   b. 75 mL of a 0.50 M C₆H₁₂O₆ solution
   c. 235 mL of a 1.80 M LiCl solution

12.120 How many grams of solute are in each of the following solutions? (12.4)
   a. 0.428 L of a 0.450 M K₂CO₃ solution
   b. 10.5 mL of a 2.50 M AgNO₃ solution
   c. 28.4 mL of a 6.00 M H₃PO₄ solution

12.121 Calculate the final concentration of the solution when water is added to prepare each of the following: (12.5)
   a. 25.0 mL of a 0.200 M NaBr solution is diluted to 50.0 mL
   b. 15.0 mL of a 12.0% (m/v) K₂SO₄ solution is diluted to 40.0 mL
   c. 75.0 mL of a 6.00 M NaOH solution is diluted to 255 mL

12.122 Calculate the final concentration of the solution when water is added to prepare each of the following: (12.5)
   a. 25.0 mL of a 18.0 M HCl solution is diluted to 500. mL
   b. 50.0 mL of a 15.0% (m/v) NH₄Cl solution is diluted to 125 mL
   c. 4.50 mL of a 8.50 M KOH solution is diluted to 75.0 mL

12.123 What is the initial volume, in milliliters, needed to prepare each of the following diluted solutions? (12.5)
   a. 250 mL of 3.0% (m/v) HCl from 10.0% (m/v) HCl
   b. 500. mL of 0.90% (m/v) NaCl from 5.0% (m/v) NaCl
   c. 350. mL of 2.00 M NaOH from 6.00 M NaOH

12.124 What is the initial volume, in milliliters, needed to prepare each of the following diluted solutions? (12.5)
   a. 250 mL of 5.0% (m/v) glucose from 20.0% (m/v) glucose
   b. 45.0 mL of 1.0% (m/v) CaCl₂ from 5.0% (m/v) CaCl₂
   c. 100. mL of 6.00 M H₂SO₄ from 18.0 M H₂SO₄

12.125 What is the final volume, in milliliters, when 25.0 mL of each of the following solutions is diluted to provide the given concentration? (12.5)
   a. 10.0% (m/v) HCl solution to give a 2.50% (m/v) HCl solution
   b. 5.00 M HCl solution to give a 1.00 M HCl solution
   c. 6.00 M HCl solution to give a 0.500 M HCl solution

12.126 What is the final volume, in milliliters, when 5.00 mL of each of the following solutions is diluted to provide the given concentration? (12.5)
   a. 20.0% (m/v) NaOH solution to give a 4.000% (m/v) NaOH solution
   b. 0.600 M NaOH solution to give a 0.100 M NaOH solution
   c. 16.0% (m/v) NaOH solution to give a 2.00% (m/v) NaOH solution

12.127 The antacid Amphojel contains aluminum hydroxide Al(OH)₃. How many milliliters of a 6.00 M HCl solution are required to react with 60.0 mL of a 2.00 M Al(OH)₃ solution? (12.6)
   \[ \text{Al(OH)}₃(s) + 3\text{HCl}(aq) \rightarrow 3\text{H}_₂\text{O}(l) + \text{AlCl}_3(aq) \]

12.128 Cadmium reacts with HCl to produce hydrogen gas and cadmium chloride. What is the molarity (M) of the HCl solution if 250. mL of the HCl solution reacts with excess cadmium to produce 4.20 L of H₂ gas measured at STP? (12.6)
   \[ \text{Cd}(s) + 2\text{HCl}(aq) \rightarrow \text{H}_₂(g) + \text{CdCl}_2(aq) \]

12.129 Calculate the freezing point of each of the following solutions: (12.7)
   a. 0.580 mol of lactose, a nonelectrolyte, added to 1.00 kg of water
   b. 45.0 g of KCl, a strong electrolyte, dissolved in 1.00 kg of water
   c. 1.5 mol of K₃PO₄, a strong electrolyte, dissolved in 1.00 kg of water

12.130 Calculate the boiling point of each of the following solutions: (12.7)
   a. 175 g of glucose, C₆H₁₂O₆, a nonelectrolyte, added to 1.00 kg of water
   b. 1.8 mol of CaCl₂, a strong electrolyte, dissolved in 1.00 kg of water
   c. 50.0 g of LiNO₃, a strong electrolyte, dissolved in 1.00 kg of water

Applications

12.131 An antacid tablet, such as Amphojel, may be taken to reduce excess stomach acid, which is 0.20 M HCl. If one dose of Amphojel contains 450 mg of Al(OH)₃, what volume, in milliliters, of stomach acid will be neutralized? (12.6)
   \[ \text{Al(OH)}₃(s) + 3\text{HCl}(aq) \rightarrow 3\text{H}_₂\text{O}(l) + \text{AlCl}_3(aq) \]

12.132 Calcium carbonate, CaCO₃, reacts with stomach acid, (HCl, hydrochloric acid) according to the following equation: (12.6)
   \[ \text{CaCO}_₃(s) + 2\text{HCl}(aq) \rightarrow \text{CO}_₂(g) + \text{H}_₂\text{O}(l) + \text{CaCl}_₂(aq) \]
   Tums, an antacid, contains CaCO₃. If Tums is added to 20.0 mL of a 0.400 M HCl solution, how many grams of CO₂ gas are produced?

12.133 Write the net ionic equation to show the formation of a solid when the following solutions are mixed. Write none if no solid forms. (12.3)
   a. K₂Cr₂O₇(aq) + BaCl₂(aq)
   b. NaOH(aq) + MgCl₂(aq)
   c. Rb₂O(aq) + Ba(NO₃)₂(aq)
   d. NaI(aq) + AgNO₃(aq)

12.134 Write the net ionic equation to show the formation of a solid when the following solutions are mixed. Write none if no solid forms. (12.3)
   a. NaBr(aq) + Hg(C₂H₃O₂)₂(aq)
   b. NH₄OH(aq) + ZnCl₂(aq)
   c. Na₂CO₃(aq) + MgBr₂(aq)
   d. NiCl₂(aq) + KBr(aq)

12.135 In a laboratory experiment, a 10.0-mL sample of NaCl solution is poured into an evaporating dish with a mass of 24.10 g. The combined mass of the evaporating dish and NaCl solution is 36.15 g. After heating, the evaporating dish and dry NaCl have a combined mass of 25.50 g. (12.4)
12.136 In a laboratory experiment, a 15.0-mL sample of KCl solution is poured into an evaporating dish with a mass of 24.10 g. The combined mass of the evaporating dish and KCl solution is 41.50 g. After heating, the evaporating dish and dry KCl have a combined mass of 28.28 g. (12.4)

a. What is the mass percent (m/m) of the KCl solution?

b. What is the total volume, in milliliters, of the solution?

c. What is the mass/volume percent (m/v) of the solution?

d. What is the molarity (M) of the solution?

12.137 Potassium fluoride has a solubility of 92 g of KF in 100 g of H₂O at 18 °C. Determine if each of the following mixtures forms an unsaturated or saturated solution at 18 °C: (12.3)

a. adding 35 g of KF to 25 g of H₂O

b. adding 42 g of KF to 50 g of H₂O

c. adding 145 g of KF to 150 g of H₂O

12.138 Copper sulfate has a solubility of 32 g of CuSO₄ in 100 g of H₂O at 20 °C. Determine if each of the following mixtures forms an unsaturated or saturated solution at 20 °C: (12.3)

a. adding 57 g of CuSO₄ to 100 g of H₂O

b. adding 24 g of CuSO₄ to 80 g of H₂O

c. adding 83 g of CuSO₄ to 150 g of H₂O

12.139 A solution is prepared with 5.0 g of HCl and 195.0 g of CuSO₄. Copper sulfate has a solubility of 32 g of CuSO₄ in 100 g of H₂O at 20 °C. Determine if each of the following mixtures forms an unsaturated or saturated solution at 20 °C: (12.3)

a. adding 57 g of CuSO₄ to 100 g of H₂O

b. adding 24 g of CuSO₄ to 80 g of H₂O

c. adding 83 g of CuSO₄ to 150 g of H₂O

12.140 A solution is prepared by dissolving 22.0 g of NaOH in 118.0 g of water. The NaOH solution has a density of 1.15 g/mL. (12.4)

a. What is the mass percent (m/m) of the NaOH solution?

b. What is the total volume, in milliliters, of the solution?

c. What is the mass/volume percent (m/v) of the solution?

d. What is the molarity (M) of the solution?
solution can be made by adding 5.00 g of glucose to enough water to make 100.0 mL of solution.

12.29 a. 17% (m/m) KCl solution
b. 5.3% (m/m) sucrose solution
c. 10% (m/m) CaCl₂ solution

12.31 a. 30% (m/v) Na₂SO₄ solution
b. 11% (m/v) sucrose solution

12.33 a. 2.5 g of KCl
b. 50. g of NH₄Cl
c. 25.0 mL of acetic acid

12.35 79.9 mL of alcohol

12.37 a. 20. g of LiNO₃ solution
b. 400. mL of KOH solution
c. 20. mL of formic acid solution

12.39 a. 0.500 M glucose solution
b. 0.0356 M KOH solution
c. 0.250 M NaCl solution

12.41 a. 120. g of NaOH
b. 59.6 g of KCl
c. 5.47 g of HCl

12.43 a. 1.50 L of KBr solution
b. 10.0 L of NaCl solution
c. 62.5 mL of Ca(NO₃)₂ solution

12.45 a. 983 mL
b. 1.70 × 10³ mL
c. 464 mL

12.47 a. 20. g of mannitol
b. 240 g of mannitol

12.49 2 L of glucose solution

12.51 Adding water (solvent) to the soup increases the volume and dilutes the tomato soup concentration.

12.53 a. 2.0 M HCl solution
b. 2.0 M NaOH solution
c. 2.5% (m/v) KOH solution
d. 3.0% (m/v) H₂SO₄ solution

12.55 a. 80. mL of HCl solution
b. 250 mL of LiCl solution
c. 600. mL of H₃PO₄ solution
d. 180 mL of glucose solution

12.57 a. 12.8 mL of 4.00 M HNO₃ solution
b. 11.9 mL of 6.00 M MgCl₂ solution
c. 1.88 mL of 8.00 M KCl solution

12.59 1.0 × 10² mL

12.61 a. 10.4 g of PbCl₂
b. 18.8 mL of Pb(NO₃)₂ solution
c. 1.20 M KCl solution

12.63 a. 206 mL of HCl solution
b. 11.2 L of H₂ gas
c. 9.09 M HCl solution

12.65 a. 74.8 mL of 3.50 M HBr solution
b. 50.4 L of H₂
c. 1.69 M HBr

12.67 a. solution
b. suspension

12.69 a. 2.0 mol of ethylene glycol in 1.0 kg of water will have a lower freezing point because it has more particles in solution.
b. 0.50 mol of MgCl₂ in 1.0 kg of water has a lower freezing point because each formula unit of MgCl₂ dissociates in water to give three particles, whereas each formula unit of KCl dissociates to give only two particles.

12.71 a. 22.3 m
b. 6.89 m

12.73 a. freezing point: -41.5 °C; boiling point: 111.6 °C
b. freezing point: -12.8 °C; boiling point: 103.6 °C

12.75 a. 10% (m/v) starch solution
b. from the 1% (m/v) starch solution into the 10% (m/v) starch solution
c. 10% (m/v) starch solution

12.77 a. B 10% (m/v) sucrose solution
b. A 8% (m/v) albumin solution
c. B 10% (m/v) starch solution

12.79 a. hypotonic
b. hypotonic
c. isotonic
d. hypertonic

12.81 a. Na⁺, Cl⁻
b. alanine
c. Na⁺, Cl⁻
d. urea

12.83 3.0 mL of chlorpromazine solution

12.85 0.50 g of CaCl₂

12.87 a. 1
b. 2
c. 1

12.89 a. 3
b. 1
c. 2

12.91 a. beaker 3
b. Na⁺ (aq) + Cl⁻ (aq) + Ag⁺ (aq) + NO₃⁻ (aq) → AgCl(s) + Na⁺ (aq) + NO₃⁻ (aq)
c. Ag⁺ (aq) + Cl⁻ (aq) → AgCl(s)

12.93 When a dried raisin is placed in a sugar solution with lower solute concentration, water flows into the cell by osmosis through the semipermeable membrane. The increase in fluid causes the raisin to swell.

12.95 a. 2
b. 1
c. 3
d. 2

12.97 The polar Cl bond leads to H-bonding with water molecules, which makes it soluble in water. Chlorine does not dissolve in hexane because it is a nonpolar solvent.

12.99 a. unsaturated solution
b. saturated solution
c. unsaturated solution

12.101 a. insoluble
b. soluble
c. insoluble
d. soluble

12.103 a. Ag⁺ (aq) + Cl⁻ (aq) → AgCl(s)

12.105 222 g of water

12.107 1.99% (m/m) ascorbic acid solution

12.109 38 mL of propyl alcohol solution

12.111 0.72 L of KOH solution

12.113 0.15 M Na₃PO₄ solution

12.115 205 mL of LiCl solution

12.117 20.5 mL of NaCl solution

12.119 a. 1600 g of Al(NO₃)₃
b. 6.8 g of C₆H₁₂O₆

c. 17.9 g of LiCl

12.121 a. 0.100 M NaBr solution
b. 4.50% (m/v) K₂SO₄ solution
c. 1.76 M NaOH solution

12.123 a. 75 mL of 10.0% (m/v) HCl solution
b. 90. mL of 5.0% (m/v) NaCl solution
c. 117 mL of 6.00 M NaOH solution

12.125 a. 100. mL
b. 125 mL
c. 300. mL
12.127 60.0 mL of HCl solution

12.129 a. −1.08 °C  
b. −2.25 °C  
c. −11 °C

12.131 87 mL of HCl solution

12.133 a. None  
b. $2\text{OH}^{-}(aq) + \text{Mg}^{2+}(aq) \rightarrow \text{Mg(OH)}_2(s)$  
c. $\text{SO}_4^{2-}(aq) + \text{Ba}^{2+}(aq) \rightarrow \text{BaSO}_4(s)$  
d. $\text{I}^- (aq) + \text{Ag}^+(aq) \rightarrow \text{AgI}(s)$

12.135 a. 11.6% (m/m) NaCl solution  
b. 2.40 M NaCl solution  
c. 0.400 M NaCl solution

12.137 a. saturated  
b. unsaturated  
c. saturated

12.139 a. 0.025% (m/m) HCl solution  
b. 134.22 mL of solution  
c. 3.73% (m/v) HCl solution  
d. 10.2 M HCl solution

12.141 0.917 M HCl solution

12.143 a. 0.60 mol of NaCl  
b. 0.30 mol of K_3PO_4

12.145 a. 2.0 m  
b. −3.7 °C
PETE R, A CHEMICAL oceanographer, is collecting data concerning the amount of dissolved gases, specifically carbon dioxide (CO$_2$) in the Atlantic Ocean. Studies indicate that CO$_2$ in the atmosphere has increased as much as 25% since the eighteenth century, which has resulted in a scientific debate regarding its effects. Peter’s research involves measuring the amount of dissolved CO$_2$ in the oceans and trying to determine its impact. The oceans are a complex mixture of many different chemicals including gases, elements and minerals, and organic and particulate matter. As a result, the oceans have been called a “chemical soup,” which can complicate a study like Peter’s. Peter understands that CO$_2$ is absorbed in the ocean through a series of equilibrium reactions. An equilibrium reaction is a reversible reaction in which both the products and the reactants are present. If the equilibrium reactions shift according to Le Châtelier’s principle, an increase in the CO$_2$ concentration could eventually increase the amount of dissolved calcium carbonate, CaCO$_3$, which makes up coral reefs and shells.

**CAREER**

**Chemical Oceanographer**

A chemical oceanographer, also called a marine chemist, studies the chemistry of the ocean. One area of study includes how chemicals or pollutants enter into and affect the ocean. These include sewage, oil or fuels, chemical fertilizers, and storm drain overflows. Oceanographers analyze how these chemicals interact with seawater, marine life, and sediments, as they can behave differently due to the ocean’s varied environmental conditions. Chemical oceanographers also study how the various elements are cycled within the ocean. For instance, oceanographers quantify the amount and rate at which carbon dioxide is absorbed at the ocean’s surface and eventually transferred to deep waters. Chemical oceanographers also aid ocean engineers in the development of instruments and vessels that enable researchers to collect data and discover previously unknown marine life.
13.1 Rates of Reactions

**LEARNING GOAL** Describe how temperature, concentration, and catalysts affect the rate of a reaction.

Earlier we looked at chemical reactions and determined the amounts of substances that react and the products that form. Now we are interested in how fast a reaction goes. If we know how fast a medication acts on the body, we can adjust the time over which the medication is taken. In construction, substances are added to cement to make it dry faster so that work can continue. Some reactions such as explosions or the formation of precipitates in a solution are very fast. When we roast a turkey or bake a cake, the reaction is slower. Some reactions such as the tarnishing of silver and the aging of the body are much slower (see Figure 13.1). We will see that some reactions need energy while other reactions produce energy. In this chapter, we will also look at the effect of changing the concentrations of reactants or products on the rate of reaction.

**FIGURE 13.1** Reaction rates vary greatly for everyday processes. A banana ripens in a few days, silver tarnishes in a few months, while the aging process of humans takes many years.

How would you compare the rates of the reaction that forms sugars in plants by photosynthesis with the reactions that digest sugars in the body?

For a chemical reaction to take place, the molecules of the reactants must come in contact with each other. The **collision theory** indicates that a reaction takes place only when molecules collide with the proper orientation and sufficient energy. Many collisions can occur, but only a few actually lead to the formation of product. For example, consider the
reaction of nitrogen and oxygen molecules (see FIGURE 13.2). To form nitrogen oxide (NO) product, the collisions between N₂ and O₂ molecules must place the atoms in the proper alignment. If the molecules are not aligned properly, no reaction takes place.

Collision that forms product

\[
\text{N}_2 + \text{O}_2 \rightarrow \text{NO} + \text{NO}
\]

Collisions that do not form product

Insufficient energy

Wrong orientation

**Activation Energy**

Even when a collision has the proper orientation, there still must be sufficient energy to break the bonds between the atoms of the reactants. The **activation energy** is the minimum amount of energy required to break the bonds between atoms of the reactants. In FIGURE 13.3, activation energy appears as an energy hill. The concept of activation energy is analogous to climbing a hill. To reach a destination on the other side, we must have the energy needed to climb to the top of the hill. Once we are at the top, we can run down the other side. The energy needed to get us from our starting point to the top of the hill would be our activation energy.

In the same way, a collision must provide enough energy to push the reactants to the top of the energy hill. Then the reactants may be converted to products. If the energy provided by the collision is less than the activation energy, the molecules simply bounce apart and no reaction occurs. The features that lead to a successful reaction are summarized next.

**FIGURE 13.2** Reacting molecules must collide, have a minimum amount of energy, and have the proper orientation to form product.

**FIGURE 13.3** The activation energy is the minimum energy needed to convert the colliding molecules into product.
Three Conditions Required for a Reaction to Occur

1. **Collision** The reactants must collide.
2. **Orientation** The reactants must align properly to break and form bonds.
3. **Energy** The collision must provide the energy of activation.

**Rate of Reaction**

The **rate** (or speed) of reaction is determined by measuring the amount of a reactant used up, or the amount of a product formed, in a certain period of time.

\[
\text{Rate of reaction} = \frac{\text{change in concentration of reactant or product}}{\text{change in time}}
\]

We can describe the rate of reaction with the analogy of eating a pizza. When we start to eat, we have a whole pizza. As time goes by, there are fewer slices of pizza left. If we know how long it took to eat the pizza, we could determine the rate at which the pizza was consumed. Let’s assume 4 slices are eaten every 8 minutes. That gives a rate of \( \frac{1}{2} \) slice per minute. After 16 minutes, all 8 slices are gone.

<table>
<thead>
<tr>
<th>Rate at Which Pizza Slices Are Eaten</th>
</tr>
</thead>
<tbody>
<tr>
<td>Slices Eaten</td>
</tr>
<tr>
<td>Time (min)</td>
</tr>
</tbody>
</table>

**Factors that Affect the Rate of a Reaction**

Reactions with low activation energies go faster than reactions with high activation energies. Some reactions go very fast, while others are very slow. For any reaction, the rate is affected by changes in temperature, changes in the concentration of the reactants, and the addition of catalysts.

**Temperature**

At higher temperatures, the increase in kinetic energy of the reactants makes them move faster and collide more often, and it provides more collisions with the required energy of activation. Reactions almost always go faster at higher temperatures. For every 10 °C increase in temperature, most reaction rates approximately double. If we want food to cook faster, we increase the temperature. When body temperature rises, there is an increase in the pulse rate, rate of breathing, and metabolic rate. If we are exposed to extreme heat, we may experience **heat stroke**, which is a condition that occurs when body temperature goes above 40.5 °C (105 °F). If the body loses its ability to regulate temperature, body temperature continues to rise, and may cause damage to the brain and internal organs. On the other hand, we slow down a reaction by decreasing the temperature. For example, we refrigerate perishable foods to make them last longer. In some cardiac surgeries, body temperature is decreased to 28 °C so the heart can be stopped and less oxygen is required by the brain. This is also the reason why some people have survived submersion in icy lakes for long periods of time. Cool water or an ice blanket may also be used to decrease the body temperature of a person with hyperthermia or heat stroke.

**Concentrations of Reactants**

For virtually all reactions, the rate of a reaction increases when the concentration of the reactants increases. When there are more reacting molecules, more collisions that form products can occur, and the reaction goes faster (see **Figure 13.4**). For example, a patient having difficulty breathing may be given a breathing mixture with a higher oxygen content than the atmosphere. The increase in the number of oxygen molecules in the lungs...
increases the rate at which oxygen combines with hemoglobin. The increased rate of oxygenation of the blood means that the patient can breathe more easily.

\[
Hb(aq) + O_2(g) \rightarrow HbO_2(aq)
\]

**Catalysts**

Another way to speed up a reaction is to lower the *energy of activation*. The energy of activation is the minimum energy needed to break apart the bonds of the reacting molecules. If a collision provides less than the activation energy, the bonds do not break and the reactant molecules bounce apart. A *catalyst* speeds up a reaction by providing an alternative pathway that has a lower energy of activation. When activation energy is lowered, more collisions provide sufficient energy for reactants to form product. During a reaction, a catalyst is not changed or consumed.

Catalysts have many uses in industry. In the manufacturing of margarine, hydrogen \((H_2)\) is added to vegetable oils. Normally, the reaction is very slow because it has a high activation energy. However, when platinum \((Pt)\) is used as a catalyst, the reaction occurs rapidly. In the body, bio-catalysts called *enzymes* make most metabolic reactions proceed at rates necessary for proper cellular activity. Enzymes are added to laundry detergents to break down proteins (proteases), starches (amylases), or greases (lipases) that have stained clothes. Such enzymes function at the low temperatures that are used in home washing machines, and they are biodegradable as well.

The factors affecting reaction rates are summarized in **TABLE 13.1**.

**SAMPLE PROBLEM 13.1 Factors That Affect the Rate of Reaction**

Indicate whether the following changes will increase, decrease, or have no effect on the rate of reaction:

a. increasing the temperature  
b. increasing the number of reacting molecules  
c. adding a catalyst

**SOLUTION**

a. A higher temperature increases the kinetic energy of the particles, which increases the number of collisions and makes more collisions effective, causing an increase in the rate of reaction.  
b. Increasing the number of reacting molecules increases the number of collisions and the rate of the reaction.  
c. Adding a catalyst increases the rate of reaction by lowering the activation energy, which increases the number of collisions that form product.

**STUDY CHECK 13.1**

How does using an ice blanket on a patient affect the rate of metabolism in the body?

**ANSWER**

Lowering the temperature will decrease the rate of metabolism.
For over 30 years, manufacturers have been required to include catalytic converters in the exhaust systems of gasoline automobile engines. When gasoline burns, the products found in the exhaust of a car contain high levels of pollutants. These include carbon monoxide (CO) from incomplete combustion, hydrocarbons such as \( \text{C}_8\text{H}_{18} \) (octane) from unburned fuel, and nitrogen oxide (NO) from the reaction of \( \text{N}_2 \) and \( \text{O}_2 \) at the high temperatures reached within the engine. Carbon monoxide is toxic, and unburned hydrocarbons and nitrogen oxide are involved in the formation of smog and acid rain.

A catalytic converter consists of solid-particle catalysts, such as platinum (Pt) and palladium (Pd), on a ceramic honeycomb that provides a large surface area and facilitates contact with pollutants. As the pollutants pass through the converter, they react with the catalysts. Today, we all use unleaded gasoline because lead interferes with the ability of the Pt and Pd catalysts in the converter to react with the pollutants. The purpose of a catalytic converter is to lower the activation energy for reactions that convert each of these pollutants into substances such as \( \text{CO}_2 \), \( \text{N}_2 \), \( \text{O}_2 \), and \( \text{H}_2\text{O} \), which are already present in the atmosphere.

\[
2\text{NO}(g) \rightarrow \text{N}_2(g) + \text{O}_2(g)
\]

NO absorbed on catalyst

\[
2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g)
\]

CO and \( \text{O}_2 \) absorbed on catalyst

\[
\text{CO}_2
\]

Surface of metal (Pt, Pd) catalyst

\[
\text{N}_2 \quad \text{O}_2
\]

NO dissociates

\[
\text{O}_2 \quad \text{CO}_2
\]

\[
\text{O}_2 \quad \text{CO}_2
\]

\[
\text{N}_2 \quad \text{O}_2
\]

O\(_2\) dissociates
13.2 Chemical Equilibrium

**LEARNING GOAL** Use the concept of reversible reactions to explain chemical equilibrium.

We consider the *forward reaction* in an equation and assumed that all of the reactants were converted to products. However, most of the time reactants are not completely converted to products because a *reverse reaction* takes place in which products collide to form the reactants. When a reaction proceeds in both a forward and reverse direction, it is said to be *reversible*. We have looked at other reversible processes. For example, the melting of solids to form liquids and the freezing of liquids to solids is a reversible physical change. Even in our daily life we have reversible events. We go from home to school and we return from school to home. We go up an escalator and we come back down. We put money in our bank account and we take money out.

An analogy for a forward and reverse reaction can be found in the phrase “We are going to the grocery store.” Although we mention our trip in one direction, we know that we will also return home from the store. Because our trip has both a forward and reverse direction, we can say the trip is reversible. It is not very likely that we would stay at the store forever.

A trip to the grocery store can be used to illustrate another aspect of reversible reactions. Perhaps the grocery store is nearby and we usually walk. However, we can change our rate. Suppose that one day we drive to the store, which increases our rate and gets us to the store faster. Correspondingly, a car also increases the rate at which we return home.

**Reversible Chemical Reactions**

A *reversible reaction* proceeds in both the forward and reverse direction. That means there are two reaction rates: one is the rate of the forward reaction, and the other is the rate of the reverse reaction. When molecules begin to react, the rate of the forward reaction is faster than the rate of the reverse reaction. As reactants are consumed and products accumulate, the rate of the forward reaction decreases and the rate of the reverse reaction increases.

**Equilibrium**

Eventually, the rates of the forward and reverse reactions become equal; the reactants form products at the same rate that the products form reactants. A reaction reaches *chemical equilibrium* when no further change takes place in the concentrations of the reactants and products, even though the two reactions continue at equal but opposite rates.
At Equilibrium:
The rate of the forward reaction is equal to the rate of the reverse reaction. No further changes occur in the concentrations of reactants and products, even though the two reactions continue at equal but opposite rates.

Let us look at the process as the reaction of \( \text{H}_2 \) and \( \text{I}_2 \) proceeds to equilibrium. Initially, only the reactants \( \text{H}_2 \) and \( \text{I}_2 \) are present. Soon, a few molecules of \( \text{HI} \) are produced by the forward reaction. With more time, additional \( \text{HI} \) molecules are produced. As the concentration of \( \text{HI} \) increases, more \( \text{HI} \) molecules collide and react in the reverse direction. As \( \text{HI} \) product builds up, the rate of the reverse reaction increases, while the rate of the forward reaction decreases. Eventually, the rates become equal, which means the reaction has reached equilibrium. Even though the concentrations remain constant at equilibrium, the forward and reverse reactions continue to occur. The forward and reverse reactions are usually shown together in a single equation by using a double arrow. A reversible reaction is two opposing reactions that occur at the same time (see Figure 13.5).

**SAMPLE PROBLEM 13.2 Reaction Rates and Equilibrium**

Complete each of the following with *equal* or *not equal*, *faster* or *slower*, *change* or *do not change*:

a. Before equilibrium is reached, the concentrations of the reactants and products ______.

b. Initially, reactants placed in a container have a _______ rate of reaction than the rate of reaction of the products.

c. At equilibrium, the rate of the forward reaction is _______ to the rate of the reverse reaction.
13.2 Chemical Equilibrium

**TRY IT FIRST**

**SOLUTION**

a. Before equilibrium is reached, the concentrations of the reactants and products *change*.  

b. Initially, reactants placed in a container have a *faster* rate of reaction than the rate of reaction of the products.  
c. At equilibrium, the rate of the forward reaction is *equal* to the rate of the reverse reaction.

**STUDY CHECK 13.2**

Complete the following statement with *change* or *do not change*:

At equilibrium, the concentrations of the reactants and products _________.

**ANSWER**

At equilibrium, the concentrations of the reactants and products *do not change*.

We can also set up a reaction starting with only reactants or with only products. Let’s look at the initial reactions in each, the forward and reverse reactions, and the equilibrium mixture that forms (see **FIGURE 13.6**).

\[
2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \iff 2\text{SO}_3(\text{g})
\]

If we start with only the reactants \(\text{SO}_2\) and \(\text{O}_2\) in the container, the reaction to form \(\text{SO}_3\) takes place until equilibrium is reached. However, if we start with only the product \(\text{SO}_3\) in the container, the reaction to form \(\text{SO}_2\) and \(\text{O}_2\) takes place until equilibrium is reached. In both containers, the equilibrium mixture contains the same concentrations of \(\text{SO}_2\), \(\text{O}_2\), and \(\text{SO}_3\).

**FIGURE 13.6** Sample (a) initially contains \(\text{SO}_2(\text{g})\) and \(\text{O}_2(\text{g})\). At equilibrium, sample (b) contains mostly \(\text{SO}_3(\text{g})\) and only small amounts of \(\text{SO}_2(\text{g})\) and \(\text{O}_2(\text{g})\), whereas sample (c) contains only \(\text{SO}_3(\text{g})\).

**Why is the same equilibrium mixture obtained from \(\text{SO}_2(\text{g})\) and \(\text{O}_2(\text{g})\) as from \(\text{SO}_3(\text{g})\)?**

**QUESTIONS AND PROBLEMS**

**13.2 Chemical Equilibrium**

**LEARNING GOAL** Use the concept of reversible reactions to explain chemical equilibrium.

**13.7** What is meant by the term reversible reaction?  
**13.8** When does a reversible reaction reach equilibrium?

**13.9** Which of the following are at equilibrium?  
a. The rate of the forward reaction is twice as fast as the rate of the reverse reaction.  
b. The concentrations of the reactants and the products do not change.  
c. The rate of the reverse reaction does not change.
13.10 Which of the following are not at equilibrium?
   a. The rates of the forward and reverse reactions are equal.
   b. The rate of the forward reaction does not change.
   c. The concentrations of reactants and the products are not constant.

13.11 The following diagrams show the chemical reaction with time:

   A ⇌ B

   If A is blue and B is orange, state whether or not the reaction has reached equilibrium in this time period and explain why.

   1 h  2 h  3 h  4 h

13.12 The following diagrams show the chemical reaction with time:

   C ⇌ D

   If C is blue and D is yellow, state whether or not the reaction has reached equilibrium in this time period and explain why.

   1 h  2 h  3 h  4 h

**13.3 Equilibrium Constants**

**LEARNING GOAL** Calculate the equilibrium constant for a reversible reaction given the concentrations of reactants and products at equilibrium.

At equilibrium, the concentrations of the reactants and products are constant. We can use a ski lift as an analogy. Early in the morning, skiers at the bottom of the mountain begin to ride the ski lift up to the slopes. After the skiers reach the top of the mountain, they ski down. Eventually, the number of people riding up the ski lift becomes equal to the number of people skiing down the mountain. There is no further change in the number of skiers on the slopes; the system is at equilibrium.

**Equilibrium Expression**

At equilibrium, the concentrations can be used to set up a relationship between the products and the reactants. Suppose we write a general equation for reactants A and B that form products C and D. The small italic letters are the coefficients in the balanced equation.

\[ aA + bB \rightleftharpoons cC + dD \]

An *equilibrium expression*, \( K_c \), for a reversible chemical reaction multiplies the concentrations of the products together and divides by the concentrations of the reactants. Each concentration is raised to a power that is equal to its coefficient in the balanced chemical equation. The square bracket around each substance indicates that the concentration is expressed in moles per liter (M). For our general reaction, this is written as:

\[ K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b} \]

We can now describe how to write the equilibrium expression for the reaction of \( \text{H}_2 \) and \( \text{I}_2 \) that forms \( \text{HI} \). The balanced chemical equation is written with a double arrow between the reactants and the products.

\[ \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g) \]
We show the concentration of the products using brackets in the numerator and the concentrations of the reactants in brackets in the denominator and write any coefficient as an exponent of its concentration (a coefficient 1 is understood).

\[
K_c = \frac{[\text{products}]}{[\text{reactants}]}
\]

**SAMPLE PROBLEM 13.3 Writing Equilibrium Expressions**

Write the equilibrium expression for the following reaction:

\[
2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)
\]

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>equation</td>
<td>equilibrium</td>
<td>expression</td>
</tr>
<tr>
<td></td>
<td>expression</td>
<td>[products]</td>
</tr>
<tr>
<td></td>
<td>[products]</td>
<td></td>
</tr>
<tr>
<td></td>
<td>[reactants]</td>
<td>[reactants]</td>
</tr>
<tr>
<td></td>
<td>[reactants]</td>
<td>[products]</td>
</tr>
</tbody>
</table>

**STEP 1** Write the balanced chemical equation.

\[
2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)
\]

**STEP 2** Write the concentrations of the products as the numerator and the reactants as the denominator.

\[
\frac{[\text{products}]}{[\text{reactants}]} = \frac{[\text{SO}_3]}{[\text{SO}_2][\text{O}_2]}
\]

**STEP 3** Write any coefficient in the equation as an exponent.

\[
K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}
\]

**STUDY CHECK 13.3**

Write the equilibrium expression for the following reaction:

\[
2\text{NO}(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}_2(g)
\]

**ANSWER**

\[
K_c = \frac{[\text{NO}_2]^2}{[\text{NO}]^2[\text{O}_2]}
\]

**Calculating Equilibrium Constants**

The **equilibrium constant**, \(K_c\), is the numerical value obtained by substituting experimentally measured molar concentrations at equilibrium into the equilibrium expression. For example, the equilibrium expression for the reaction of \(\text{H}_2\) and \(\text{I}_2\) is written

\[
\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g) \quad K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}
\]

In the first experiment, the molar concentrations for the reactants and products at equilibrium are found to be \([\text{H}_2] = 0.10 \text{ M}, [\text{I}_2] = 0.20 \text{ M},\) and \([\text{HI}] = 1.04 \text{ M}.\) When we substitute these values into the equilibrium expression, we obtain the numerical value of \(K_c\).

In additional experiments 2 and 3, the mixtures have different equilibrium concentrations for the system at equilibrium at the same temperature. However, when these concentrations are used to calculate the equilibrium constant, we obtain the same value of \(K_c\) for each (see **TABLE 13.2**). Thus, a reaction at a specific temperature can have only one value for the equilibrium constant.
The units of $K_c$ depend on the specific equation. In this example, the units of $[M]^2/[M]^2$ cancel out to give a value of 54. In other equations, the concentration units do not cancel. However, in this text, the numerical value will be given without any units as shown in Sample Problem 13.4.

### TABLE 13.2 Equilibrium Constant for $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ at 427 °C

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Concentrations at Equilibrium</th>
<th>Equilibrium Constant</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>$[H_2]$</td>
<td>$[I_2]$</td>
</tr>
<tr>
<td>1</td>
<td>0.10 M</td>
<td>0.20 M</td>
</tr>
<tr>
<td>2</td>
<td>0.20 M</td>
<td>0.20 M</td>
</tr>
<tr>
<td>3</td>
<td>0.30 M</td>
<td>0.17 M</td>
</tr>
</tbody>
</table>

### SAMPLE PROBLEM 13.4 Calculating an Equilibrium Constant

The decomposition of dinitrogen tetroxide forms nitrogen dioxide.

\[ N_2O_4(g) \rightleftharpoons 2NO_2(g) \]

What is the numerical value of $K_c$ at 100 °C if a reaction mixture at equilibrium contains 0.45 M $N_2O_4$ and 0.31 M $NO_2$?

#### TRY IT FIRST

#### SOLUTION

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.45 M $N_2O_4$</td>
<td>$K_c$</td>
<td>equilibrium expression</td>
</tr>
<tr>
<td>0.31 M $NO_2$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Equation**

\[ N_2O_4(g) \rightleftharpoons 2NO_2(g) \]

**STEP 2** Write the equilibrium expression, $K_c$.

\[ K_c = \frac{[NO_2]^2}{[N_2O_4]} \]

**STEP 3** Substitute equilibrium (molar) concentrations and calculate $K_c$.

\[ K_c = \frac{[0.31]^2}{[0.45]} = 0.21 \]

**STUDY CHECK 13.4**

Calculate the numerical value of $K_c$ if an equilibrium mixture contains 0.040 M $NH_3$, 0.60 M $H_2$, and 0.20 M $N_2$.

\[ 2NH_3(g) \rightleftharpoons 3H_2(g) + N_2(g) \]

**ANSWER**

$K_c = 27$
Heterogeneous Equilibrium

Up to now, our examples have been reactions that involve only gases. A reaction in which all the reactants and products are in the same state reaches **homogeneous equilibrium**. When the reactants and products are in two or more states, the equilibrium is termed a **heterogeneous equilibrium**. In the following reaction, solid calcium carbonate reaches heterogeneous equilibrium with solid calcium oxide and carbon dioxide gas (see **FIGURE 13.7**).

\[
\text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g)
\]

In contrast to gases, the concentrations of pure solids and pure liquids are constant; they do not change. Therefore, pure solids and liquids are not included in the equilibrium expression. For this heterogeneous equilibrium, the \( K_c \) expression does not include the concentration of \( \text{CaCO}_3(s) \) or \( \text{CaO}(s) \). It is written as \( K_c = [\text{CO}_2] \).

**SAMPLE PROBLEM 13.5 Heterogeneous Equilibrium Expression**

Write the equilibrium expression for the following reaction at equilibrium:

\[
4\text{HCl}(g) + \text{O}_2(g) \rightleftharpoons 2\text{H}_2\text{O}(l) + 2\text{Cl}_2(g)
\]

**TRY IT FIRST**

**SOLUTION**

**STEP 1** Write the balanced chemical equation.

\[
4\text{HCl}(g) + \text{O}_2(g) \rightleftharpoons 2\text{H}_2\text{O}(l) + 2\text{Cl}_2(g)
\]

**STEP 2** Write the concentrations of the products as the numerator and the reactants as the denominator. In this heterogeneous reaction, the concentration of the \( \text{H}_2\text{O} \), which is a pure liquid, is not included in the equilibrium expression.

\[
\frac{[\text{Products}]}{[\text{Reactants}]} = \frac{[\text{Cl}_2]^2}{[\text{HCl}][\text{O}_2]}
\]

**STEP 3** Write any coefficient in the equation as an exponent.

\[
K_c = \frac{[\text{Cl}_2]^2}{[\text{HCl}]^4[\text{O}_2]}
\]

**STUDY CHECK 13.5**

Solid iron(II) oxide and carbon monoxide gas react to produce solid iron and carbon dioxide gas. Write the balanced chemical equation and the equilibrium expression for this reaction at equilibrium.

**ANSWER**

\[
\text{FeO}(s) + \text{CO}(g) \rightleftharpoons \text{Fe}(s) + \text{CO}_2(g)
\]

\[
K_c = \frac{[\text{CO}_2]}{[\text{CO}]}
\]

**QUESTIONS AND PROBLEMS**

**13.3 Equilibrium Constants**

**LEARNING GOAL** Calculate the equilibrium constant for a reversible reaction given the concentrations of reactants and products at equilibrium.

13.13 Write the equilibrium expression for each of the following reactions:

a. \( \text{CH}_4(g) + 2\text{H}_2\text{S}(g) \rightleftharpoons \text{CS}_2(g) + 4\text{H}_2(g) \)

b. \( 2\text{NO}(g) \rightleftharpoons \text{N}_2(g) + \text{O}_2(g) \)

c. \( 2\text{SO}_3(g) \rightleftharpoons \text{SO}_2(g) + 4\text{O}_2(g) \)

d. \( \text{CH}_4(g) + \text{H}_2\text{O}(g) \rightleftharpoons 3\text{H}_2(g) + \text{CO}(g) \)

13.14 Write the equilibrium expression for each of the following reactions:

a. \( 2\text{HBr}(g) \rightleftharpoons \text{H}_2(g) + \text{Br}_2(g) \)

b. \( 2\text{BrNO}(g) \rightleftharpoons \text{Br}_2(g) + 2\text{NO}(g) \)

c. \( \text{CH}_4(g) + \text{Cl}_2(g) \rightleftharpoons \text{CH}_2\text{Cl}(g) + \text{HCl}(g) \)

d. \( \text{Br}_2(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{BrCl}(g) \)
13.15 Write the equilibrium expression for the reaction in the diagram and calculate the numerical value of $K_c$. In the diagram, X atoms are orange and Y atoms are blue.

$$X_3(g) + Y_2(g) \rightleftharpoons 2XY(g)$$

[Diagram of X and Y atoms]

13.16 Write the equilibrium expression for the reaction in the diagram and calculate the numerical value of $K_c$. In the diagram, A atoms are red and B atoms are green.

$$2AB(g) \rightleftharpoons A_2(g) + B_2(g)$$

[Diagram showing red and green atoms]

13.17 What is the numerical value of $K_c$ for the following reaction if the equilibrium mixture contains 0.30 M N$_2$O$_4$ and 0.21 M NO$_2$?

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

13.18 What is the numerical value of $K_c$ for the following reaction if the equilibrium mixture contains 0.30 M CO$_2$, 0.033 M H$_2$, 0.20 M CO, and 0.30 M H$_2$O?

$$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$$

13.19 What is the numerical value of $K_c$ for the following reaction if the equilibrium mixture contains 0.51 M CO, 0.30 M H$_2$, 1.8 M CH$_4$, and 2.0 M H$_2$O?

$$CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$$

13.20 What is the numerical value of $K_c$ for the following reaction if the equilibrium mixture contains 0.44 M N$_2$, 0.40 M H$_2$, and 2.2 M NH$_3$?

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

13.21 Identify each of the following as a homogeneous or heterogeneous equilibrium:

a. $2O_3(g) \rightleftharpoons 3O_2(g)$
b. $2NaHCO_3(s) \rightleftharpoons Na_2CO_3(s) + CO_2(g) + H_2O(g)$
c. $C_6H_6(g) + 3H_2(g) \rightleftharpoons C_6H_12(g)$
d. $4HCl(g) + Si(l) \rightleftharpoons SiCl_4(l) + 2H_2(g)$

13.22 Identify each of the following as a homogeneous or heterogeneous equilibrium:

a. $CO(g) + H_2(g) \rightleftharpoons C(s) + H_2O(g)$
b. $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$
c. $CS_2(g) + 4H_2(g) \rightleftharpoons CH_4(g) + 2H_2S(g)$
d. $Ti(s) + 2Cl_2(g) \rightleftharpoons TiCl_4(g)$

13.23 Write the equilibrium expression for each of the reactions in problem 13.21.

13.24 Write the equilibrium expression for each of the reactions in problem 13.22.

13.25 What is the numerical value of $K_c$ for the following reaction if the equilibrium mixture at 750 °C contains 0.20 M CO and 0.052 M CO$_2$?

$$FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g)$$

13.26 What is the numerical value of $K_c$ for the following reaction if the equilibrium mixture at 800 °C contains 0.030 M CO$_2$?

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

### 13.4 Using Equilibrium Constants

#### LEARNING GOAL

Use an equilibrium constant to predict the extent of reaction and to calculate equilibrium concentrations.

The values of $K_c$ can be large or small. The size of the equilibrium constant depends on whether equilibrium is reached with more products than reactants, or more reactants than products. However, the size of an equilibrium constant does not affect how fast equilibrium is reached.

#### Equilibrium with a Large $K_c$

When a reaction has a large equilibrium constant, it means that the forward reaction produced a large amount of products when equilibrium was reached. Then the equilibrium mixture contains mostly products, which makes the concentrations of the products in the numerator higher than the concentrations of the reactants in the denominator. Thus at equilibrium, this reaction has a large $K_c$. Consider the reaction of SO$_2$ and O$_2$, which has a large $K_c$. At equilibrium, the reaction mixture contains mostly product and few reactants (see **FIGURE 13.8**).

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

**Mostly product**

**Few reactants**

$$3.4 \times 10^2$$
13.4 Using Equilibrium Constants

Equilibrium with a Small \( K_c \)

When a reaction has a small equilibrium constant, the equilibrium mixture contains a high concentration of reactants and a low concentration of products. Then the equilibrium expression has a small number in the numerator and a large number in the denominator. Thus at equilibrium, this reaction has a small \( K_c \).

\[
N_2(g) + O_2(g) \rightleftharpoons 2NO(g)
\]

\[
K_c = \frac{[NO]^2}{[N_2][O_2]} \quad \text{Few products} \quad \text{Mostly reactants} = 2 \times 10^{-9}
\]

A few reactions have equilibrium constants close to 1, which means they have about equal concentrations of reactants and products. Moderate amounts of reactants have been converted to products upon reaching equilibrium (see Figure 13.10).

**FIGURE 13.8**  In the reaction of \( SO_2(g) \) and \( O_2(g) \), the equilibrium mixture contains mostly product \( SO_3(g) \), which results in a large \( K_c \).

Why does an equilibrium mixture containing mostly product have a large \( K_c \)?

**FIGURE 13.9**  The equilibrium mixture contains a very small amount of the product NO and a large amount of the reactants \( N_2 \) and \( O_2 \), which results in a small \( K_c \).

Does a reaction with a \( K_c = 3.2 \times 10^{-5} \) contain mostly reactants or products at equilibrium?

**FIGURE 13.10**  At equilibrium, a reaction with a large \( K_c \) contains mostly products, whereas a reaction with a small \( K_c \) contains mostly reactants.

Does a reaction with a \( K_c = 1.2 \) contain mostly reactants, mostly products, or about equal amounts of both reactants and products at equilibrium?
TABLE 13.3 lists some equilibrium constants and the extent of their reaction.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
<th>$K_c$</th>
<th>Equilibrium Mixture Contains</th>
</tr>
</thead>
<tbody>
<tr>
<td>$2\text{CO}(g) + \text{O}_2(g) \rightleftharpoons 2\text{CO}_2(g)$</td>
<td>$2 \times 10^{11}$</td>
<td>Mostly product</td>
<td></td>
</tr>
<tr>
<td>$2\text{H}_2(g) + \text{S}_2(g) \rightleftharpoons 2\text{H}_2\text{S}(g)$</td>
<td>$1.1 \times 10^7$</td>
<td>Mostly product</td>
<td></td>
</tr>
<tr>
<td>$\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g)$</td>
<td>$1.6 \times 10^3$</td>
<td>Mostly product</td>
<td></td>
</tr>
<tr>
<td>$\text{PCl}_3(g) \rightleftharpoons \text{PCl}_5(g) + \text{Cl}_2(g)$</td>
<td>$1.2 \times 10^{-2}$</td>
<td>Mostly reactant</td>
<td></td>
</tr>
<tr>
<td>$\text{N}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}(g)$</td>
<td>$2 \times 10^{-9}$</td>
<td>Mostly reactants</td>
<td></td>
</tr>
</tbody>
</table>

Calculating Concentrations at Equilibrium

When we know the numerical value of the equilibrium constant and all the equilibrium concentrations except one, we can calculate the unknown concentration as shown in Sample Problem 13.6.

**SAMPLE PROBLEM 13.6 Calculating Concentration Using an Equilibrium Constant**

For the reaction of carbon dioxide and hydrogen, the equilibrium concentrations are 0.25 M $\text{CO}_2$, 0.80 M $\text{H}_2$, and 0.50 M $\text{H}_2\text{O}$. What is the equilibrium concentration of $\text{CO}(g)$?

$\text{CO}_2(g) + \text{H}_2(g) \rightleftharpoons \text{CO}(g) + \text{H}_2\text{O}(g)$ $K_c = 0.11$

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>Equation</td>
<td>$\text{CO}_2(g) + \text{H}_2(g) \rightleftharpoons \text{CO}(g) + \text{H}_2\text{O}(g)$</td>
<td>$K_c = 0.11$</td>
<td></td>
</tr>
<tr>
<td>$0.25 \text{ M CO}_2$, $0.80 \text{ M H}_2$, $0.50 \text{ M H}_2\text{O}$</td>
<td>$\text{[CO]}$</td>
<td>$K_c$ expression</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write the equilibrium expression, $K_c$, and solve for the needed concentration.

$$K_c = \frac{[\text{CO}][\text{H}_2\text{O}]}{[\text{CO}_2][\text{H}_2]}$$

We rearrange $K_c$ to solve for the unknown $[\text{CO}]$ as follows:

Multiply both sides by $[\text{CO}_2][\text{H}_2]$.

$$K_c \times [\text{CO}_2][\text{H}_2] = \frac{[\text{CO}][\text{H}_2\text{O}]}{[\text{CO}_2][\text{H}_2]} \times [\text{CO}_2][\text{H}_2]$$

Divide both sides by $[\text{H}_2\text{O}]$.

$$K_c \times \frac{[\text{CO}_2][\text{H}_2]}{[\text{H}_2\text{O}]} = \frac{[\text{CO}][\text{H}_2\text{O}]}{[\text{H}_2\text{O}]}$$

$$[\text{CO}] = K_c \times \frac{[\text{CO}_2][\text{H}_2]}{[\text{H}_2\text{O}]}$$
13.4 Using Equilibrium Constants

STEP 3 Substitute the equilibrium (molar) concentrations and calculate the needed concentration.

\[
[\text{CO}] = K_c \times \frac{[\text{CO}_2][\text{H}_2]}{[\text{H}_2\text{O}]} = 0.11 \times \frac{[0.25][0.80]}{[0.50]} = 0.044 \text{ M}
\]

STUDY CHECK 13.6
When ethene (C₂H₄) reacts with water vapor, ethanol (C₂H₆O) is produced. If an equilibrium mixture contains 0.020 M C₂H₄ and 0.015 M H₂O, what is the equilibrium concentration of C₂H₆O? At 327 °C, the \( K_c \) is \( 9.0 \times 10^3 \).

\[
\text{C}_2\text{H}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{C}_2\text{H}_6\text{O}(\text{g})
\]

ANSWER

\([\text{C}_2\text{H}_6\text{O}] = 2.7 \ \text{M}\)

QUESTIONS AND PROBLEMS

13.4 Using Equilibrium Constants

LEARNING GOAL Use an equilibrium constant to predict the extent of reaction and to calculate equilibrium concentrations.

13.27 If the \( K_c \) for this reaction is 4, which of the following diagrams represents the molecules in an equilibrium mixture? In the diagrams, X atoms are orange and Y atoms are blue.

\[
\text{X}_2(\text{g}) + \text{Y}_2(\text{g}) \rightleftharpoons 2\text{XY}(\text{g})
\]

A B C

13.28 If the \( K_c \) for this reaction is 2, which of the following diagrams represents the molecules in an equilibrium mixture? In the diagrams, A atoms are orange and B atoms are green.

\[
2\text{AB}(\text{g}) \rightleftharpoons \text{A}_2(\text{g}) + \text{B}_2(\text{g})
\]

A B C

13.29 Indicate whether each of the following equilibrium mixtures contains mostly products or mostly reactants:

a. \( \text{Cl}_2(\text{g}) + \text{NO}(\text{g}) \rightleftharpoons 2\text{NOCl}(\text{g}) \) \( K_c = 3.7 \times 10^8 \)

b. \( 2\text{H}_2(\text{g}) + \text{S}_2(\text{g}) \rightleftharpoons 2\text{H}_2\text{S}(\text{g}) \) \( K_c = 1.1 \times 10^7 \)

c. \( 3\text{O}_2(\text{g}) \rightleftharpoons 2\text{O}_3(\text{g}) \) \( K_c = 1.7 \times 10^{-56} \)

13.30 Indicate whether each of the following equilibrium mixtures contains mostly products or mostly reactants:

a. \( \text{CO}(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{COCl}_2(\text{g}) \) \( K_c = 5.0 \times 10^{-9} \)

b. \( 2\text{HF}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{F}_2(\text{g}) \) \( K_c = 1.0 \times 10^{-95} \)

c. \( 2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g}) \) \( K_c = 6.0 \times 10^{13} \)

13.31 The equilibrium constant, \( K_c \), for this reaction is 54.

\[
\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})
\]

If the equilibrium mixture contains 0.015 M I₂ and 0.030 M HI, what is the molar concentration of H₂?

13.32 The equilibrium constant, \( K_c \), for the following reaction is \( 4.6 \times 10^{-3} \). If the equilibrium mixture contains 0.050 M NO₂, what is the molar concentration of N₂O₄?

\[
\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})
\]

13.33 The \( K_c \) for the following reaction at 100 °C is 2.0. If the equilibrium mixture contains 2.0 M NO and 1.0 M Br₂, what is the molar concentration of NOBr?

\[
2\text{NOBr}(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{Br}_2(\text{g})
\]

13.34 The \( K_c \) for the following reaction at 225 °C is \( 1.7 \times 10^2 \). If the equilibrium mixture contains 0.18 M H₂ and 0.020 M N₂, what is the molar concentration of NH₃?

\[
3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})
\]
13.5 Changing Equilibrium Conditions: Le Châtelier’s Principle

**LEARNING GOAL.** Use Le Châtelier’s principle to describe the changes made in equilibrium concentrations when reaction conditions change.

We have seen that when a reaction reaches equilibrium, the rates of the forward and reverse reactions are equal and the concentrations remain constant. Now we will look at what happens to a system at equilibrium when changes occur in reaction conditions, such as changes in concentration, volume, and temperature.

**Le Châtelier’s Principle**

When we alter any of the conditions of a system at equilibrium, the rates of the forward and reverse reactions will no longer be equal. We say that a stress is placed on the equilibrium. Then the system responds by changing the rate of the forward or reverse reaction in the direction that relieves that stress to reestablish equilibrium. We can use **Le Châtelier’s principle**, which states that when a system at equilibrium is disturbed, the system will shift in the direction that will reduce that stress.

**Le Châtelier’s Principle**

When a stress (change in conditions) is placed on a reaction at equilibrium, the equilibrium will shift in the direction that relieves the stress.

Suppose we have two water tanks connected by a pipe. When the water levels in the tanks are equal, water flows in the forward direction from Tank A to Tank B at the same rate as it flows in the reverse direction from Tank B to Tank A. Suppose we add more water to Tank A. With a higher level of water in Tank A, more water flows in the forward direction from Tank A to Tank B than in the reverse direction from Tank B to Tank A, which is shown with a longer arrow. Eventually, equilibrium is reached as the levels in both tanks become equal, but higher than before. Then the rate of water flows equally between Tank A and Tank B.

![Diagram of water tanks](image)

**Effect of Concentration Changes on Equilibrium**

We will now use the reaction of $\text{H}_2$ and $\text{I}_2$ to illustrate how a change in concentration disturbs the equilibrium and how the system responds to that stress.

$$\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g)$$

Suppose that more of the reactant $\text{H}_2$ is added to the equilibrium mixture, which increases the concentration of $\text{H}_2$. Because $K_c$ cannot change for a reaction at a given
temperature, adding more $H_2$ places a stress on the system (see FIGURE 13.11).

Then the system relieves this stress by increasing the rate of the forward reaction, which is indicated by the direction of the large arrow. Thus, more product is formed until the system is again at equilibrium. According to Le Châtelier’s principle, adding more reactant causes the system to shift in the direction of the product until equilibrium is reestablished.

\[
\text{Add } H_2 \\
H_2(g) + I_2(g) \quad \overset{\text{Forward}}{\longrightarrow} \quad 2 HI(g)
\]

Suppose now that some $H_2$ is removed from the reaction mixture at equilibrium, which lowers the concentration of $H_2$ and slows the rate of the forward reaction. Using Le Châtelier’s principle, we know that when some of the reactants are removed, the system will shift in the direction of the reactants until equilibrium is reestablished.

\[
\text{Remove } H_2 \\
H_2(g) + I_2(g) \quad \overset{\text{Reverse}}{\longrightarrow} \quad 2 HI(g)
\]

The concentrations of the products of an equilibrium mixture can also increase or decrease. For example, if more HI is added, there is an increase in the rate of the reaction in the reverse direction, which converts some of the product to reactants. The concentration of the products decreases and the concentration of the reactants increases until equilibrium is reestablished. Using Le Châtelier’s principle, we see that the addition of a product causes the system to shift in the direction of the reactants.

\[
\text{Add HI} \\
H_2(g) + I_2(g) \quad \overset{\text{Reverse}}{\longrightarrow} \quad 2 HI(g)
\]

In another example, some HI is removed from an equilibrium mixture, which decreases the concentration of the product. Then there is a shift in the direction of the product to reestablish equilibrium.

\[
\text{Remove HI} \\
H_2(g) + I_2(g) \quad \overset{\text{Forward}}{\longrightarrow} \quad 2 HI(g)
\]

In summary, Le Châtelier’s principle indicates that a stress caused by adding a substance at equilibrium is relieved when the equilibrium system shifts the reaction...
away from that substance. Adding more reactant causes an increase in the forward reaction to products. Adding more products causes an increase in the reverse reaction to reactants. When some of a substance is removed, the equilibrium system shifts in the direction of that substance. These features of Le Châtelier’s principle are summarized in Table 13.4.

### Effect of a Catalyst on Equilibrium

Sometimes a catalyst is added to a reaction to speed up a reaction by lowering the activation energy. As a result, the rates of both the forward and reverse reactions increase. The time required to reach equilibrium is shorter, but the same ratios of products and reactants are attained. Therefore, a catalyst speeds up the forward and reverse reactions, but it has no effect on the equilibrium mixture.

### Effect of Volume Change on Equilibrium

If there is a change in the volume of a gas mixture at equilibrium, there will also be a change in the concentrations of those gases. Decreasing the volume will increase the concentration of gases, whereas increasing the volume will decrease their concentration. Then the system responds to reestablish equilibrium.

Let’s look at the effect of decreasing the volume of the equilibrium mixture of the following reaction:

\[ 2\text{CO}(g) + \text{O}_2(g) \rightleftharpoons 2\text{CO}_2(g) \]

If we decrease the volume, all the concentrations increase. According to Le Châtelier’s principle, the increase in concentration is relieved when the system shifts in the direction of the smaller number of moles.

On the other hand, when the volume of the equilibrium gas mixture increases, the concentrations of all the gases decrease. Then the system shifts in the direction of the greater number of moles to reestablish equilibrium (see Figure 13.12).

When a reaction has the same number of moles of reactants as products, a volume change does not affect the equilibrium mixture because the concentrations of the reactants and products change in the same way.

\[ \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g) \]


\[ \begin{array}{c|c}
\text{Stress} & \text{Shift in the Direction of} \\
\hline
\text{Increasing [H}_2\text{]} & \text{Product} \\
\text{Decreasing [H}_2\text{]} & \text{Reactants} \\
\text{Increasing [I}_2\text{]} & \text{Product} \\
\text{Decreasing [I}_2\text{]} & \text{Reactants} \\
\text{Increasing [HI]} & \text{Reactants} \\
\text{Decreasing [HI]} & \text{Product} \\
\end{array} \]

### Table 13.4 Effect of Concentration Changes on Equilibrium

\[ \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g) \]
The transport of oxygen involves an equilibrium between hemoglobin (Hb), oxygen, and oxyhemoglobin (HbO₂).

\[
\text{Hb}(aq) + \text{O}_2(g) \rightleftharpoons \text{HbO}_2(aq)
\]

When the \( \text{O}_2 \) level is high in the alveoli of the lung, the reaction shifts in the direction of the product \( \text{HbO}_2 \). In the tissues where \( \text{O}_2 \) concentration is low, the reverse reaction releases the oxygen from the hemoglobin. The equilibrium expression is written

\[
K_c = \frac{[\text{HbO}_2]}{[\text{Hb}][\text{O}_2]}
\]

At normal atmospheric pressure, oxygen diffuses into the blood because the partial pressure of oxygen in the alveoli is higher than that in the blood. At an altitude above 8000 ft, a decrease in the atmospheric pressure results in a significant reduction in the partial pressure of oxygen, which means that less oxygen is available for the blood and body tissues. The fall in atmospheric pressure at higher altitudes decreases the partial pressure of inhaled oxygen, and there is less driving pressure for gas exchange in the lungs. At an altitude of 18 000 ft, a person will obtain 29% less oxygen. When oxygen levels are lowered, a person may experience hypoxia, characterized by increased respiratory rate, headache, decreased mental acuteness, fatigue, decreased physical coordination, nausea, vomiting, and cyanosis. A similar problem occurs in persons with a history of lung disease that impairs gas diffusion in the alveoli or in persons with a reduced number of red blood cells, such as smokers.

According to Le Châtelier’s principle, we see that a decrease in oxygen will shift the equilibrium in the direction of the reactants. Such a shift depletes the concentration of \( \text{HbO}_2 \) and causes the hypoxia.

\[
\text{Hb}(aq) + \text{O}_2(g) \rightleftharpoons \text{HbO}_2(aq)
\]

Immediate treatment of altitude sickness includes hydration, rest, and if necessary, descending to a lower altitude. The adaptation to lowered oxygen levels requires about 10 days. During this time the bone marrow increases red blood cell production, providing more red blood cells and more hemoglobin. A person living at a high altitude can have 50% more red blood cells than someone at sea level. This increase in hemoglobin causes a shift in the equilibrium back in the direction of \( \text{HbO}_2 \) product. Eventually, the higher concentration of \( \text{HbO}_2 \) will provide more oxygen to the tissues and the symptoms of hypoxia will lessen.

For some who climb high mountains, it is important to stop and acclimatize for several days at increasing altitudes. At very high altitudes, it may be necessary to use an oxygen tank.

**Effect of a Change in Temperature on Equilibrium**

We can think of heat as a reactant or a product in a reaction. For example, in the equation for an endothermic reaction, heat is written on the reactant side. When the temperature of an endothermic reaction increases, the system responds by shifting in the direction of the products to remove heat.

\[
\text{N}_2(g) + \text{O}_2(g) + \text{heat} \rightleftharpoons 2\text{NO}(g)
\]

If the temperature is decreased for an endothermic reaction, there is a decrease in heat. Then the system shifts in the direction of the reactants to add heat.

\[
\text{N}_2(g) + \text{O}_2(g) + \text{heat} \rightleftharpoons 2\text{NO}(g)
\]

In the equation for an exothermic reaction, heat is written on the product side. When the temperature of an exothermic reaction increases, the system responds by shifting in the direction of the reactants to remove heat.

\[
2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g) + \text{heat}
\]
If the temperature is decreased for an exothermic reaction, there is a decrease in heat. Then the system shifts in the direction of the products to add heat.

\[
2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g}) + \text{heat}
\]

**TABLE 13.5** summarizes the ways we can use Le Châtelier’s principle to determine the shift in equilibrium that relieves a stress caused by the change in a condition.

**TABLE 13.5** Effects of Condition Changes on Equilibrium

<table>
<thead>
<tr>
<th>Condition</th>
<th>Change (Stress)</th>
<th>Shift in the Direction of</th>
</tr>
</thead>
<tbody>
<tr>
<td>Concentration</td>
<td>Adding a reactant</td>
<td>Products (forward reaction)</td>
</tr>
<tr>
<td></td>
<td>Removing a reactant</td>
<td>Reactants (reverse reaction)</td>
</tr>
<tr>
<td></td>
<td>Adding a product</td>
<td>Reactants (reverse reaction)</td>
</tr>
<tr>
<td></td>
<td>Removing a product</td>
<td>Products (forward reaction)</td>
</tr>
<tr>
<td>Volume (container)</td>
<td>Decreasing the volume</td>
<td>Fewer moles of gas</td>
</tr>
<tr>
<td></td>
<td>Increasing the volume</td>
<td>More moles of gas</td>
</tr>
<tr>
<td>Temperature</td>
<td><strong>Endothermic reaction</strong></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Increasing the temperature</td>
<td>Products (forward reaction to remove heat)</td>
</tr>
<tr>
<td></td>
<td>Decreasing the temperature</td>
<td>Reactants (reverse reaction to add heat)</td>
</tr>
<tr>
<td></td>
<td><strong>Exothermic reaction</strong></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Increasing the temperature</td>
<td>Reactants (reverse reaction to remove heat)</td>
</tr>
<tr>
<td></td>
<td>Decreasing the temperature</td>
<td>Products (forward reaction to add heat)</td>
</tr>
<tr>
<td>Catalyst</td>
<td>Increasing the rates equally</td>
<td>No effect</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 13.7** Using Le Châtelier’s Principle

Methanol, \(\text{CH}_3\text{O}\), is finding use as a fuel additive. Describe the effect of each of the following changes on the equilibrium mixture for the combustion of methanol:

\[
2\text{CH}_3\text{O}(\text{g}) + 3\text{O}_2(\text{g}) \rightleftharpoons 2\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g}) + 1450 \text{kJ}
\]

a. adding more \(\text{CO}_2\)
b. adding more \(\text{O}_2\)
c. increasing the volume of the container
d. increasing the temperature
e. adding a catalyst

**TRY IT FIRST**

**SOLUTION**

a. When the concentration of the product \(\text{CO}_2\) increases, the equilibrium shifts in the direction of the reactants.
b. When the concentration of the reactant \(\text{O}_2\) increases, the equilibrium shifts in the direction of the products.
c. When the volume increases, the equilibrium shifts in the direction of the greater number of moles of gas, which is the products.
d. When the temperature is increased for an exothermic reaction, the equilibrium shifts in the direction of the reactants to remove heat.
e. When a catalyst is added, there is no change in the equilibrium mixture.
STUDY CHECK 13.7
Describe the effect of each of the following changes on the equilibrium mixture for the following reaction:

\[ 2HF(g) + Cl_2(g) + 357 \text{ kJ} \rightleftharpoons 2HCl(g) + F_2(g) \]

a. adding more Cl_2  
b. decreasing the volume of the container  
c. decreasing the temperature

ANSWER

a. When the concentration of the reactant Cl_2 increases, the equilibrium shifts in the direction of the products.  
b. There is no change in equilibrium mixture because the moles of reactants are equal to the moles of products.  
c. When the temperature for an endothermic reaction decreases, the equilibrium shifts to add heat, which is in the direction of the reactants.

CHEMISTRY LINK TO HEALTH

Homeostasis: Regulation of Body Temperature

In a physiological system of equilibrium called homeostasis, changes in our environment are balanced by changes in our bodies. It is crucial to our survival that we balance heat gain with heat loss. If we do not lose enough heat, our body temperature rises. At high temperatures, the body can no longer regulate our metabolic reactions. If we lose too much heat, body temperature drops. At low temperatures, essential functions proceed too slowly.

The skin plays an important role in the maintenance of body temperature. When the outside temperature rises, receptors in the skin send signals to the brain. The temperature-regulating part of the brain stimulates the sweat glands to produce perspiration. As perspiration evaporates from the skin, heat is removed and the body temperature is decreased.

In cold temperatures, epinephrine is released, causing an increase in metabolic rate, which increases the production of heat. Receptors on the skin signal the brain to constrict the blood vessels. Less blood flows through the skin, and heat is conserved. The production of perspiration stops, thereby lessening the heat lost by evaporation.

Questions and Problems

13.5 Changing Equilibrium Conditions: Le Châtelier’s Principle

**LEARNING GOAL** Use Le Châtelier’s principle to describe the changes made in equilibrium concentrations when reaction conditions change.

13.35 In the lower atmosphere, oxygen is converted to ozone (O₃) by the energy provided from lightning.

\[ 3O_2(g) + \text{heat} \rightleftharpoons 2O_3(g) \]

For each of the following changes at equilibrium, indicate whether the equilibrium shifts in the direction of product, reactants, or does not change:

a. adding more O₂(g)  
b. adding more O₃(g)  
c. increasing the temperature  
d. increasing the volume of the container  
e. adding a catalyst

13.36 Ammonia is produced by reacting nitrogen gas and hydrogen gas.

\[ N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + 92 \text{ kJ} \]

For each of the following changes at equilibrium, indicate whether the equilibrium shifts in the direction of product, reactants, or does not change:

a. removing some N₂(g)  
b. decreasing the temperature  
c. adding more NH₃(g)  
d. adding more H₂(g)  
e. increasing the volume of the container
13.37 Hydrogen chloride can be made by reacting hydrogen gas and chlorine gas.

\[ \text{H}_2(g) + \text{Cl}_2(g) + \text{heat} \rightleftharpoons 2\text{HCl}(g) \]

For each of the following changes at equilibrium, indicate whether the equilibrium shifts in the direction of product, reactants, or does not change:

a. adding more \( \text{H}_2(g) \)

b. increasing the temperature

c. removing some \( \text{HCl}(g) \)

d. adding a catalyst

13.38 When heated, carbon monoxide reacts with water to produce carbon dioxide and hydrogen.

\[ \text{CO}(g) + \text{H}_2\text{O}(g) \rightleftharpoons \text{CO}_2(g) + \text{H}_2(g) + \text{heat} \]

For each of the following changes at equilibrium, indicate whether the equilibrium shifts in the direction of products, reactants, or does not change:

a. decreasing the temperature

b. adding more \( \text{H}_2(g) \)

c. removing \( \text{CO}_2(g) \) as its forms

d. adding more \( \text{H}_2\text{O}(g) \)

e. decreasing the volume of the container

### Applications

Use the following equation for the equilibrium of hemoglobin in the blood to answer problems 13.39 and 13.40:

\[ \text{Hb}(aq) + \text{O}_2(g) \rightleftharpoons \text{HbO}_2(aq) \]

13.39 Athletes who train at high altitudes initially experience hypoxia, which causes their body to produce more hemoglobin. When an athlete first arrives at high altitude,

a. What is the stress on the hemoglobin equilibrium?

b. In what direction does the hemoglobin equilibrium shift?

c. What is the stress on the hemoglobin equilibrium?

d. In what direction does the hemoglobin equilibrium shift?

13.40 A person who has been a smoker and has a low oxygen blood saturation uses an oxygen tank for supplemental oxygen. When oxygen is first supplied,

a. What is the stress on the hemoglobin equilibrium?

b. In what direction does the hemoglobin equilibrium shift?

### 13.6 Equilibrium in Saturated Solutions

#### LEARNING GOAL

Write the solubility product expression for a slightly soluble ionic compound and calculate the \( K_{sp} \); use \( K_{sp} \) to determine the solubility.

Until now, we have looked primarily at equilibrium systems that involve gases. However, there are also equilibrium systems that involve aqueous saturated solutions between solid solutes of slightly soluble ionic compounds and their ions. Everyday examples of solubility equilibrium in solution are found in the slightly soluble ionic compounds that are found in bone and kidney stones. Bone is composed of calcium phosphate, \( \text{Ca}_3(\text{PO}_4)_2 \), which produces ions during bone loss. Kidney stones are composed of compounds such as calcium oxalate, \( \text{CaC}_2\text{O}_4 \).

### Solubility Product Expression

In a saturated solution, a solid slightly soluble ionic compound is in equilibrium with its ions. As long as the temperature remains constant, the concentration of the ions in the saturated solution is constant. Let us look at the solubility equilibrium equation for \( \text{CaC}_2\text{O}_4 \), which is written with the solid solute on the left and the ions in solution on the right.

\[ \text{CaC}_2\text{O}_4(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{C}_2\text{O}_4^{2-}(aq) \]

The solubility of a substance is the quantity that dissolves to form a saturated solution. We represent the solubility in a saturated aqueous solution of solid \( \text{CaC}_2\text{O}_4 \) by the **solubility product expression** \( (K_{sp}) \), which is the product of the ion concentrations. As in other heterogeneous equilibria, the concentration of the solid \( \text{CaC}_2\text{O}_4 \) is constant and is not included in the solubility product expression.

\[ K_{sp} = [\text{Ca}^{2+}][\text{C}_2\text{O}_4^{2-}] \]

In another example, we look at the equilibrium of solid calcium phosphate and its ions \( \text{Ca}^{2+} \) and \( \text{PO}_4^{3-} \).

\[ \text{Ca}_3(\text{PO}_4)_2(s) \rightleftharpoons 3\text{Ca}^{2+}(aq) + 2\text{PO}_4^{3-}(aq) \]

As with other equilibrium expressions, the molar concentration of each product ion is raised to a power that is equal to its coefficient in the balanced equilibrium equation. For this equilibrium equation, the solubility product expression consists of \( [\text{Ca}^{2+}] \) raised to the power of 3 and \( [\text{PO}_4^{3-}] \) raised to the power of 2.

\[ K_{sp} = [\text{Ca}^{2+}]^3[\text{PO}_4^{3-}]^2 \]
Solubility Product Constant

The numerical value of the solubility product expression is the solubility product constant, \( K_{sp} \). In this text, solubility will be expressed as the molar solubility, which is the moles of solute that dissolve in 1 liter of saturated solution. The calculation of a solubility product constant is shown in Sample Problem 13.8.

### SAMPLE PROBLEM 13.8 Calculating the Solubility Product Constant

We can make a saturated solution of BaCO₃ by adding solid BaCO₃ to water and stirring until equilibrium is reached. What is the numerical value of \( K_{sp} \) for BaCO₃ if the equilibrium mixture contains \( 5.1 \times 10^{-5} \text{ M} \) Ba²⁺ and \( 5.1 \times 10^{-5} \text{ M} \) CO₃²⁻?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>([\text{Ba}^{2+}] = 5.1 \times 10^{-5} \text{ M}, \quad [\text{CO}_3^{2-}] = 5.1 \times 10^{-5} \text{ M} )</td>
<td>numerical solubility product value of ( K_{sp} ) expression</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write the equilibrium equation for the dissociation of the slightly soluble ionic compound.

\[ \text{BaCO}_3(s) \rightleftharpoons \text{Ba}^{2+}(aq) + \text{CO}_3^{2-}(aq) \]

**STEP 3** Write the solubility product expression, \( K_{sp} \).

\[ K_{sp} = [\text{Ba}^{2+}][\text{CO}_3^{2-}] \]

**STEP 4** Substitute the molar concentration of each ion into the \( K_{sp} \) expression and calculate.

\[ K_{sp} = [5.1 \times 10^{-5}][5.1 \times 10^{-5}] = 2.6 \times 10^{-9} \]

**STUDY CHECK 13.8**

A saturated solution of AgBr contains \( 7.3 \times 10^{-7} \text{ M} \) Ag⁺ and \( 7.3 \times 10^{-7} \text{ M} \) Br⁻. What is the numerical value of \( K_{sp} \) for AgBr?

**ANSWER**

\[ K_{sp} = 5.3 \times 10^{-13} \]

**TABLE 13.6** gives values of \( K_{sp} \) for a selected group of slightly soluble ionic compounds at 25 °C.

### SAMPLE PROBLEM 13.9 Calculating the Solubility Product Constant Using Coefficients

A saturated solution of strontium fluoride, SrF₂, contains \( 8.7 \times 10^{-4} \text{ M} \) Sr²⁺ and \( 1.7 \times 10^{-3} \text{ M} \) F⁻. What is the numerical value of \( K_{sp} \) for SrF₂?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>([\text{Sr}^{2+}] = 8.7 \times 10^{-4} \text{ M}, \quad [\text{F}^-] = 1.7 \times 10^{-3} \text{ M} )</td>
<td>numerical solubility product value of ( K_{sp} ) expression</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

---

**Guide to Calculating \( K_{sp} \)**

**STEP 1** State the given and needed quantities.

**STEP 2** Write the equilibrium equation for the dissociation of the slightly soluble ionic compound.

**STEP 3** Write the solubility product expression, \( K_{sp} \).

**STEP 4** Substitute the molar concentration of each ion into the \( K_{sp} \) expression and calculate.

**TABLE 13.6** Examples of Solubility Product Constants (\( K_{sp} \))

<table>
<thead>
<tr>
<th>Formula</th>
<th>( K_{sp} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>AgCl</td>
<td>( 1.8 \times 10^{-10} )</td>
</tr>
<tr>
<td>Ag₂SO₄</td>
<td>( 1.2 \times 10^{-5} )</td>
</tr>
<tr>
<td>BaSO₄</td>
<td>( 1.1 \times 10^{-10} )</td>
</tr>
<tr>
<td>CaCO₃</td>
<td>( 5.0 \times 10^{-9} )</td>
</tr>
<tr>
<td>CaF₂</td>
<td>( 3.2 \times 10^{-11} )</td>
</tr>
<tr>
<td>Ca(OH)₂</td>
<td>( 6.5 \times 10^{-6} )</td>
</tr>
<tr>
<td>CaSO₄</td>
<td>( 2.4 \times 10^{-5} )</td>
</tr>
<tr>
<td>PbCl₂</td>
<td>( 1.5 \times 10^{-6} )</td>
</tr>
<tr>
<td>PbCO₃</td>
<td>( 7.4 \times 10^{-14} )</td>
</tr>
</tbody>
</table>

Terraces in Pamukkale, Turkey, are composed of calcium carbonate, which is slightly soluble.
STEP 2 Write the equilibrium equation for the dissociation of the slightly soluble ionic compound.

\[ \text{SrF}_2(s) \rightleftharpoons \text{Sr}^{2+}(aq) + 2\text{F}^-(aq) \]

STEP 3 Write the solubility product expression, \( K_{sp} \).

\[ K_{sp} = [\text{Sr}^{2+}][\text{F}^-]^2 \]

STEP 4 Substitute the molar concentration of each ion into the \( K_{sp} \) expression and calculate.

\[ K_{sp} = [8.7 \times 10^{-3}][1.7 \times 10^{-3}]^2 = 2.5 \times 10^{-9} \]

**STUDY CHECK 13.9**

What is the numerical value of \( K_{sp} \) for silver oxalate, \( \text{Ag}_2\text{C}_2\text{O}_4 \), if a saturated solution contains \( 2.2 \times 10^{-4} \text{ M} \) \( \text{Ag}^+ \) and \( 1.1 \times 10^{-4} \text{ M} \) \( \text{C}_2\text{O}_4^{2-} \)?

**ANSWER**

\[ K_{sp} = 5.3 \times 10^{-12} \]

---

**Molar Solubility, \( S \)**

The molar solubility, \( S \), of a slightly soluble ionic compound is the number of moles of solute that dissolves in 1 liter of solution. For example, the molar solubility of cadmium sulfide, \( \text{CdS} \), is found experimentally to be \( 1 \times 10^{-12} \text{ M} \).

\[ \text{CdS}(s) \rightleftharpoons \text{Cd}^{2+}(aq) + \text{S}^{2-}(aq) \]

Because \( \text{CdS} \) dissociates into \( \text{Cd}^{2+} \) and \( \text{S}^{2-} \) ions, they each have a concentration equal to the solubility \( S \).

\[ S = [\text{Cd}^{2+}] = [\text{S}^{2-}] = 1 \times 10^{-12} \text{ M} \]

If we know the \( K_{sp} \) of a slightly soluble ionic compound, we can determine its molar solubility, as shown in Sample Problem 13.10.

---

**SAMPLE PROBLEM 13.10 Calculating the Molar Solubility from \( K_{sp} \)**

Calculate the molar solubility, \( S \), of \( \text{PbSO}_4 \) if it has a \( K_{sp} = 1.6 \times 10^{-8} \).

**TRY IT FIRST**

**SOLUTION**

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>( K_{sp} = 1.6 \times 10^{-8} )</td>
<td>molar solubility (( S )) of ( \text{PbSO}_4(s) )</td>
<td>solubility product expression</td>
<td></td>
</tr>
</tbody>
</table>

STEP 2 Write the equilibrium equation for the dissociation of the slightly soluble ionic compound.

\[ \text{PbSO}_4(s) \rightleftharpoons \text{Pb}^{2+}(aq) + \text{SO}_4^{2-}(aq) \]

STEP 3 Write the solubility product expression, \( K_{sp} \), using \( S \).

\[ K_{sp} = [\text{Pb}^{2+}][\text{SO}_4^{2-}] = S \times S = S^2 = 1.6 \times 10^{-8} \]
**STEP 4** Calculate the molar solubility, $S$.

\[ S^2 = 1.6 \times 10^{-8} \]
\[ S = \sqrt{1.6 \times 10^{-8}} = 1.3 \times 10^{-4} \text{ M} \]

Thus, $1.3 \times 10^{-4}$ mol of PbSO$_4$ will dissolve in 1 L of solution.

**STUDY CHECK 13.10**

Calculate the molar solubility, $S$, of MnS if it has a $K_{sp} = 2.5 \times 10^{-10}$.

**ANSWER**

\[ S = 1.6 \times 10^{-5} \text{ M} \]

---

**QUESTIONS AND PROBLEMS**

**13.6 Equilibrium in Saturated Solutions**

**LEARNING GOAL** Write the solubility product expression for a slightly soluble ionic compound and calculate $K_{sp}$; use $K_{sp}$ to determine the solubility.

**13.41** For each of the following slightly soluble ionic compounds, write the equilibrium equation for dissociation and the solubility product expression:

- **a.** MgCO$_3$
- **b.** CaF$_2$
- **c.** Ag$_3$PO$_4$

**13.42** For each of the following slightly soluble ionic compounds, write the equilibrium equation for dissociation and the solubility product expression:

- **a.** Ag$_2$S
- **b.** Al(OH)$_3$
- **c.** BaF$_2$

**13.43** A saturated solution of barium sulfate, BaSO$_4$, has $[\text{Ba}^{2+}] = 1 \times 10^{-5}$ M and $[\text{SO}_4^{2-}] = 1 \times 10^{-5}$ M. What is the numerical value of $K_{sp}$ for BaSO$_4$?

**13.44** A saturated solution of copper(II) sulfide, CuS, has $[\text{Cu}^{2+}] = 1.1 \times 10^{-18}$ M and $[\text{S}^{2-}] = 1.1 \times 10^{-18}$ M. What is the numerical value of $K_{sp}$ for CuS?

**13.45** A saturated solution of silver carbonate, Ag$_2$CO$_3$, has $[\text{Ag}^+] = 2.6 \times 10^{-4}$ M and $[\text{CO}_3^{2-}] = 1.3 \times 10^{-4}$ M. What is the numerical value of $K_{sp}$ for Ag$_2$CO$_3$?

**13.46** A saturated solution of barium fluoride, BaF$_2$, has $[\text{Ba}^{2+}] = 3.6 \times 10^{-3}$ M and $[\text{F}^-] = 7.2 \times 10^{-3}$ M. What is the numerical value of $K_{sp}$ for BaF$_2$?

**13.47** Calculate the molar solubility, $S$, of CuI if it has a $K_{sp}$ of $1 \times 10^{-12}$.

**13.48** Calculate the molar solubility, $S$, of SnS if it has a $K_{sp}$ of $1 \times 10^{-26}$.

---

**Follow Up**

**EQUILIBRIUM OF CO$_2$ IN THE OCEAN**

One of Peter’s research projects was to study the influence of a change in ocean water acidity on coral species, which have calcium carbonate skeletons. The acidity is affected by the concentration of carbon dioxide dissolved in the water. The carbon dioxide dissolves readily in water to form H$_2$CO$_3$(aq), which breaks up into H$^+$(aq) and HCO$_3^-$ (aq). The increase in H$^+$ makes the ocean water more acidic.

**Applications**

**13.49**

- **a.** Write the equation for the dissociation of the slightly soluble ionic compound calcium carbonate.
- **b.** Write the solubility product expression, $K_{sp}$.

**13.50**

- **a.** What is the numerical value of $K_{sp}$ if the equilibrium mixture has $[\text{Ca}^{2+}]$ and $[\text{CO}_3^{2-}]$ equal to $7.1 \times 10^{-5}$ M?
- **b.** If increasing the acidity of the ocean has the effect of decreasing the concentration of CO$_3^{2-}$, will calcium carbonate be more or less soluble if the acidity increases?
CHAPTER REVIEW

13.1 Rates of Reactions

**LEARNING GOAL** Describe how temperature, concentration, and catalysts affect the rate of a reaction.

- The rate of a reaction is the speed at which the reactants are converted to products.
- Increasing the concentrations of reactants, increasing the temperature, or adding a catalyst can increase the rate of a reaction.

13.2 Chemical Equilibrium

**LEARNING GOAL** Use the concept of reversible reactions to explain chemical equilibrium.

- Chemical equilibrium occurs in a reversible reaction when the rate of the forward reaction becomes equal to the rate of the reverse reaction.
- At equilibrium, no further change occurs in the concentrations of the reactants and products as the forward and reverse reactions continue.

13.3 Equilibrium Constants

**LEARNING GOAL** Calculate the equilibrium constant for a reversible reaction given the concentrations of reactants and products at equilibrium.

- An equilibrium constant, $K_c$, is the ratio of the concentrations of the products to the concentrations of the reactants, with each concentration raised to a power equal to its coefficient in the balanced chemical equation.
- For heterogeneous reactions, only the molar concentrations of gases are placed in the equilibrium expression.

13.4 Using Equilibrium Constants

**LEARNING GOAL** Use an equilibrium constant to predict the extent of reaction and to calculate equilibrium concentrations.
• A large value of $K_c$ indicates that an equilibrium mixture contains mostly products and few reactants, whereas a small value of $K_c$ indicates that the equilibrium mixture contains mostly reactants.
• Equilibrium constants can be used to calculate the concentration of a component in the equilibrium mixture.

### 13.5 Changing Equilibrium Conditions: Le Châtelier’s Principle

**LEARNING GOAL** Use Le Châtelier’s principle to describe the changes made in equilibrium concentrations when reaction conditions change.

- When reactants are removed or products are added to an equilibrium mixture, the system shifts in the direction of the reactants.
- When reactants are added or products are removed from an equilibrium mixture, the system shifts in the direction of the products.
- A decrease in the volume of a reaction container causes a shift in the direction of the smaller number of moles of gas.

### 13.6 Equilibrium in Saturated Solutions

**LEARNING GOAL** Write the solubility product expression for a slightly soluble ionic compound and calculate $K_{sp}$, use $K_{sp}$ to determine the solubility.

- In a saturated solution, a slightly soluble ionic compound is in equilibrium with its ions.
- In a saturated solution, the concentrations of the ions from the slightly soluble ionic compound are constant and can be used to calculate the solubility product constant, $K_{sp}$.
- If $K_{sp}$ for a slightly soluble ionic compound is known, its solubility can be calculated.

### KEY TERMS

- **activation energy** The energy that must be provided by a collision to break apart the bonds of the reacting molecules.
- **catalyst** A substance that increases the rate of reaction by lowering the activation energy.
- **chemical equilibrium** The point at which the rate of forward and reverse reactions are equal so that no further change in concentrations of reactants and products takes place.
- **collision theory** A model for a chemical reaction stating that molecules must collide with sufficient energy and proper orientation to form products.
- **equilibrium constant, $K_c$** The numerical value obtained by substituting the equilibrium concentrations of the components into the equilibrium expression.
- **equilibrium expression** The ratio of the concentrations of products to the concentrations of reactants, with each component raised to an exponent equal to the coefficient of that compound in the balanced chemical equation.

### CORE CHEMISTRY SKILLS

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

#### Writing the Equilibrium Expression (13.3)

- An equilibrium expression for a reversible reaction is written by multiplying the concentrations of the products in the numerator and dividing by the product of the concentrations of the reactants in the denominator.
- Each concentration is raised to a power equal to its coefficient in the balanced chemical equation:

$$K_c = \frac{[\text{Products}]}{[\text{Reactants}]} = \frac{[C]^c [D]^d}{[A]^e [B]^f}$$

**Example:** Write the equilibrium expression for the following chemical reaction:

$$2\text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g)$$

**Answer:**

$$K_c = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2}$$

#### Calculating an Equilibrium Constant (13.3)

- The equilibrium constant, $K_c$, is the numerical value obtained by substituting experimentally measured molar concentrations at equilibrium into the equilibrium expression.

**Example:** Calculate the numerical value of $K_c$ for the following reaction when the equilibrium mixture contains 0.025 M NO$_2$ and 0.087 M N$_2$O$_4$:

$$2\text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g)$$
Answer: Write the equilibrium expression, substitute the molar concentrations, and calculate.

\[ K_c = \frac{[N_2O_4]}{[NO]^2} = \frac{[0.087]}{[0.025]^2} = 140 \]

Calculating Equilibrium Concentrations (13.4)

- To determine the concentration of a product or a reactant at equilibrium, we use the equilibrium expression to solve for the unknown concentration.

Example: Calculate the equilibrium concentration for CF₄ if \( K_c = 2.0 \), and the equilibrium mixture contains 0.10 M COF₂ and 0.050 M CO₂.

\[ 2\text{COF}_2(g) \rightleftharpoons \text{CO}_2(g) + \text{CF}_4(g) \]

Answer: Write the equilibrium expression.

\[ K_c = \frac{[\text{CO}_2][\text{CF}_4]}{[\text{COF}_2]^2} \]

Rearrange the equation, substitute the molar concentrations, and calculate.

\[ [\text{CF}_4] = K_c \times \frac{[\text{COF}_2]^2}{[\text{CO}_2]} = 2.0 \times \frac{(0.10)^2}{0.050} = 0.40 \text{ M} \]

Using Le Châtelier’s Principle (13.5)

- Le Châtelier’s principle states that when a system at equilibrium is disturbed by changes in concentration, volume, or temperature, the system will shift in the direction that will reduce that stress.

Example: Nitrogen and oxygen form dinitrogen pentoxide in an exothermic reaction.

\[ 2\text{N}_2(g) + 5\text{O}_2(g) \rightleftharpoons 2\text{N}_2\text{O}_5(g) + \text{heat} \]

For each of the following changes at equilibrium, indicate whether the equilibrium shifts in the direction of products, reactants, or does not change:

a. removing some \( \text{N}_2(g) \)
b. decreasing the temperature
c. increasing the volume of the container

Answer: a. Removing a reactant shifts the equilibrium in the direction of the reactants.
b. Decreasing the temperature shifts the equilibrium of an exothermic reaction in the direction of products, which produces heat.
c. Increasing the volume of the container shifts the equilibrium in the direction of the greater number of moles of gas, which is in the direction of the reactants.

Writing the Solubility Product Expression (13.6)

- In a saturated solution, a solid slightly soluble ionic compound is in equilibrium with its ions.

\[ \text{FeF}_2(s) \rightleftharpoons \text{Fe}^{2+}(aq) + 2\text{F}^-(aq) \]

We represent the solubility of solid \( \text{FeF}_2 \) in a saturated aqueous solution by the solubility product expression, \( K_{sp} \), which is the product of the ion concentrations raised to the powers equal to the coefficients.

\[ K_{sp} = [\text{Fe}^{2+}][\text{F}^-]^2 \]

Example: Write the equilibrium equation for the dissociation of the slightly soluble ionic compound \( \text{PbI}_2 \), and its solubility product expression.

Answer: \( \text{PbI}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{I}^-(aq) \) \( K_{sp} = [\text{Pb}^{2+}][\text{I}^-]^2 \)

Calculating a Solubility Product Constant (13.6)

- The solubility product constant, \( K_{sp} \), is calculated by substituting the molar concentrations of the ions into the solubility product expression.

Example: What is the numerical value of \( K_{sp} \) for a saturated solution of \( \text{PbI}_2 \) that contains \( 1.3 \times 10^{-3} \text{ M} \) \( \text{Pb}^{2+} \) and \( 2.6 \times 10^{-3} \text{ M} \) \( \text{I}^- \)?

Answer: The \( K_{sp} \) is calculated by substituting the molar concentrations into the solubility product expression.

\[ K_{sp} = [\text{Pb}^{2+}][\text{I}^-]^2 = (1.3 \times 10^{-3})(2.6 \times 10^{-3})^2 = 8.8 \times 10^{-9} \]

Calculating the Molar Solubility (13.6)

- The molar solubility, \( S \), of a slightly soluble ionic compound is the number of moles of solute that dissolves in 1 liter of solution.
- If we know the \( K_{sp} \) of a slightly soluble ionic compound, we can calculate the molar solubility, \( S \), of the slightly soluble ionic compound.

Example: What is the solubility product expression and the molar solubility, \( S \), of \( \text{NiS} \), which has a \( K_{sp} \) of \( 4.0 \times 10^{-20} \)?

Answer: \( K_{sp} = [\text{Ni}^{2+}][\text{S}^2^-] = S \times S = S^2 = 4.0 \times 10^{-20} \)

\[ S = \sqrt{4.0 \times 10^{-20}} = 2.0 \times 10^{-10} \text{ M} \]

Therefore, \( 2.0 \times 10^{-10} \) mol of \( \text{NiS} \) will dissolve in 1 L of saturated solution.

UNDErstanding THE CoNCEPTS

The chapter sections to review are shown in parentheses at the end of each question.

13.51 Write the equilibrium constant expression for each of the following reactions: (13.3)

a. \( \text{CS}_2(g) + 3\text{Cl}_2(g) \rightleftharpoons \text{S}_2\text{Cl}_2(g) + \text{CCl}_4(g) \)
b. \( 2\text{NO}(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{NOCl}(g) \)
c. \( 5\text{CO}(g) + 1\text{O}_2(s) \rightleftharpoons \text{I}_2(g) + 5\text{CO}_2(g) \)

13.52 Write the equilibrium constant expression for each of the following reactions: (13.3)

a. \( 2\text{N}_2\text{O}(g) + \text{O}_2(g) \rightleftharpoons 4\text{NO}(g) \)
b. \( \text{COCl}_2(g) \rightleftharpoons \text{CO}(g) + \text{Cl}_2(g) \)
c. \( \text{Ni}(s) + 4\text{CO}(g) \rightleftharpoons \text{Ni}(_4\text{CO})_4(g) \)

13.53 Would the equilibrium constant, \( K_c \), for the reaction in the diagrams have a large or small value? (13.4)

Initially

At equilibrium
13.54 Would the equilibrium constant, \( K_c \), for the reaction in the diagrams have a large or small value? (13.4)

\[
\begin{array}{ccc}
\begin{array}{c}
\text{Initially} \quad \text{At equilibrium}
\end{array} & \\
\text{+} & \text{+} & \text{+} & \text{+}
\end{array}
\]

13.55 Would \( T_2 \) be higher or lower than \( T_1 \) for the reaction shown in the diagrams? (13.5)

\[
\begin{array}{ccc}
\begin{array}{c}
T_1 = 300 \, ^\circ C
\end{array} & \\
\text{+} & \text{+} & \text{+} & \text{+} & \text{heat}
\end{array}
\]

13.56 Would the reaction shown in the diagrams be exothermic or endothermic? (13.5)

\[
\begin{array}{ccc}
\begin{array}{c}
T_1 = 100 \, ^\circ C
\end{array} & \\
\text{+} & \text{+} & \text{+} & \text{+}
\end{array}
\]

\[
\begin{array}{ccc}
\begin{array}{c}
T_2 = ?
\end{array} & \\
\text{+} & \text{+} & \text{+} & \text{+}
\end{array}
\]

### ADDITIONAL QUESTIONS AND PROBLEMS

13.57 For each of the following changes at equilibrium, indicate whether the equilibrium shifts in the direction of products, reactants, or does not change: (13.5)

\[
\begin{align*}
\text{C(s)} + \text{H}_2\text{O(g)} + \text{heat} & \rightleftharpoons \text{CO(g)} + \text{H}_2(g) \\
a. \text{increasing the temperature of the reaction} & \\
b. \text{decreasing the volume of the reaction container} & \\
c. \text{adding a catalyst} & \\
d. \text{adding more} \text{H}_2\text{O(g)}
\end{align*}
\]

13.58 For each of the following changes at equilibrium, indicate whether the equilibrium shifts in the direction of products, reactants, or does not change: (13.5)

\[
\begin{align*}
\text{N}_2(g) + \text{O}_2(g) + \text{heat} & \rightleftharpoons 2\text{NO(g)} \\
a. \text{increasing the temperature} & \\
b. \text{decreasing the volume of the container} & \\
c. \text{adding a catalyst} & \\
d. \text{adding more} \text{N}_2(g)
\end{align*}
\]

13.59 For each of the following reactions, indicate if the equilibrium mixture contains mostly products, mostly reactants, or both reactants and products: (13.4)

\[
\begin{align*}
a. \text{N}_2(g) + \text{O}_2(g) & \rightleftharpoons 2\text{NO(g)} \quad K_c = 5 \times 10^{-31} \\
b. \text{2CO(g)} + \text{O}_2(g) & \rightleftharpoons 2\text{CO}_2(g) \quad K_c = 2 \times 10^{11} \\
c. \text{PCl}_3(g) & \rightleftharpoons \text{PCl}_5(g) + \text{Cl}_2(g) \quad K_c = 1.2 \times 10^{-2} \\
d. \text{H}_2(g) + \text{F}_2(g) & \rightleftharpoons 2\text{HF(g)} \quad K_c = 1.15 \times 10^2
\end{align*}
\]

13.60 For each of the following reactions, indicate if the equilibrium mixture contains mostly products, mostly reactants, or both reactants and products: (13.4)

\[
\begin{align*}
a. \text{2H}_2\text{O(g)} & \rightleftharpoons \text{2H}_2(g) + \text{O}_2(g) \quad K_c = 4 \times 10^{-48} \\
b. \text{N}_2(g) + 3\text{H}_2(g) & \rightleftharpoons 2\text{NH}_3(g) \quad K_c = 0.30 \\
c. \text{2SO}_2(g) + \text{O}_2(g) & \rightleftharpoons 2\text{SO}_3(g) \quad K_c = 1.2 \times 10^9 \\
d. \text{H}_2(g) + \text{S(s)} & \rightleftharpoons 2\text{H}_2\text{S(g)} \quad K_c = 7.8 \times 10^5
\end{align*}
\]

13.61 Consider the reaction: (13.3)

\[
\text{COCl}_2(g) \rightleftharpoons \text{CO(g)} + \text{Cl}_2(g)
\]

\[
\begin{align*}
a. \text{Write the equilibrium constant expression.} & \\
b. \text{What is the numerical value of} K_c \text{ for the reaction if the concentrations at equilibrium are 0.50} M \text{ COCl}_2, 5.0 M \text{ CO}, \text{ and 0.25} M \text{ Cl}_2? &
\end{align*}
\]

13.62 Consider the reaction: (13.3)

\[
\text{2SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)
\]

\[
\begin{align*}
a. \text{Write the equilibrium expression.} & \\
b. \text{What is the numerical value of} K_c \text{ for the reaction if the concentrations at equilibrium are 0.10} M \text{ SO}_2, 0.12 M \text{ O}_2, \text{ and 0.60} M \text{ SO}_3? &
\end{align*}
\]

13.63 The \( K_c \) for the following reaction is 10.0 at 110 °C. If an equilibrium mixture contains 1.0 M \( \text{O}_2 \), what is the molar concentration of \( \text{O}_3 \)? (13.3, 13.4)

\[
\text{3O}_2(g) \rightleftharpoons 2\text{O}_3(g)
\]

13.64 The \( K_c \) for the following reaction is 15 at 220 °C. If an equilibrium mixture contains 0.40 M \( \text{CO} \) and 0.20 M \( \text{H}_2 \), what is the molar concentration of \( \text{CH}_3\text{O} \)? (13.3, 13.4)

\[
\text{CO(g)} + 2\text{H}_2(g) \rightleftharpoons \text{CH}_3\text{O(g)}
\]

13.65 According to Le Châtelier’s principle, does the equilibrium shift in the direction of products or reactants when \( \text{O}_2 \) is added to the equilibrium mixture of each of the following reactions? (13.5)

\[
\begin{align*}
a. \text{3O}_2(g) & \rightleftharpoons 2\text{O}_3(g) \\
b. \text{2CO}_2(g) & \rightleftharpoons 2\text{CO(g)} + \text{O}_2(g) \\
c. \text{2SO}_2(g) + \text{O}_2(g) & \rightleftharpoons 2\text{SO}_3(g) \\
d. \text{2SO}_2(g) + 2\text{H}_2\text{O(g)} & \rightleftharpoons 2\text{H}_2\text{S(g)} + 3\text{O}_2(g)
\end{align*}
\]
13.66 According to Le Châtelier’s principle, does the equilibrium shift in the direction of products or reactants when N₂ is added to the equilibrium mixture of each of the following reactions? (13.5)

a. 2NH₃(g) ⇌ 3H₂(g) + N₂(g)  
   b. N₂(g) + O₂(g) ⇌ 2NO(g)  
   c. 2NO₂(g) ⇌ N₂(g) + 2O₂(g)  
   d. 4NH₃(g) + 3O₂(g) ⇌ 2N₂(g) + 6H₂O(g)

13.67 Would decreasing the volume of the container for each of the following reactions cause the equilibrium to shift in the direction of products, reactants, or not change? (13.5)

a. 3O₂(g) ⇌ 2O₃(g)  
   b. 2CO₂(g) ⇌ 2CO(g) + O₂(g)  
   c. P₄(g) + 5O₂(g) ⇌ P₂O₁₀(s)  
   d. 2SO₂(g) + 2H₂O(g) ⇌ 2H₂S(g) + 3O₂(g)

13.68 Would increasing the volume of the container for each of the following reactions cause the equilibrium to shift in the direction of products, reactants, or not change? (13.5)

a. 2NH₃(g) ⇌ 3H₂(g) + N₂(g)  
   b. N₂(g) + O₂(g) ⇌ 2NO(g)  
   c. N₂(g) + 2O₂(g) ⇌ 2NO₂(g)  
   d. 4NH₃(g) + 3O₂(g) ⇌ 2N₂(g) + 6H₂O(g)

13.69 The equilibrium constant, Kₑ, for the decomposition of COCl₂ to CO and Cl₂ is 0.68. If an equilibrium mixture contains 0.40 M CO and 0.74 M Cl₂, what is the molar concentration of COCl₂? (13.3, 13.4)

$$
\text{COCl}_2(g) \rightleftharpoons \text{CO}(g) + \text{Cl}_2(g)
$$

13.70 The equilibrium constant, Kₑ, for the reaction of carbon and water to form carbon monoxide and hydrogen is 0.20 at 1000 °C. If an equilibrium mixture contains solid carbon, 0.40 M H₂O, and 0.40 M CO, what is the molar concentration of H₂? (13.3, 13.4)

$$
\text{C(s)} + \text{H}_2\text{O}(g) \rightleftharpoons \text{CO}(g) + \text{H}_2(g)
$$

13.71 For each of the following slightly soluble ionic compounds, write the equilibrium equation for dissociation and the solubility product expression: (13.6)

a. KF  
   b. Mg(OH)₂  
   c. Ca₃(PO₄)₂

13.72 For each of the following slightly soluble ionic compounds, write the equilibrium equation for dissociation and the solubility product expression: (13.6)

a. NiS  
   b. PbBr₂  
   c. Al(OH)₃

13.73 An unknown ionic compound has the formula A₂X₂ and is known to be slightly soluble in water. The saturated solution of A₂X₂ has [A⁺²] = 4.3 × 10⁻⁸ M and [X⁻] = 8.3 × 10⁻³ M. What is the numerical value of Kₛ for A₂X₂? (13.6)

13.74 A saturated solution of copper(I) chloride, CuCl, has [Cu⁺] = 1.1 × 10⁻⁶ M and [Cl⁻] = 1.1 × 10⁻³ M. What is the numerical value of Kₛ for CuCl? (13.6)

13.75 A saturated solution of manganese(II) hydroxide, Mn(OH)₂, has [Mn²⁺] = 3.7 × 10⁻⁵ M and [OH⁻] = 7.4 × 10⁻³ M. What is the numerical value of Kₛ for Mn(OH)₂? (13.6)

13.76 A saturated solution of silver chromate, Ag₂CrO₄, has [Ag⁺] = 1.3 × 10⁻⁸ M and [CrO₄²⁻] = 6.5 × 10⁻⁸ M. What is the numerical value of Kₛ for Ag₂CrO₄? (13.6)

13.77 What is the molar solubility, S, of CdS if it has a Kₛ of 1.0 × 10⁻²⁰? (13.6)

13.78 What is the molar solubility, S, of CuCO₃ if it has a Kₛ of 1.0 × 10⁻¹⁰? (13.6)

**CHALLENGE QUESTIONS**

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

13.79 The Kₑ at 100 °C is 2.0 for the decomposition reaction of NOBr. (13.3, 13.4, 13.5)

$$
2\text{NOBr}(g) \rightleftharpoons 2\text{NO}(g) + \text{Br}_2(g)
$$

In an experiment, 1.0 mol of NOBr, 1.0 mol of NO, and 1.0 mol of Br₂ were placed in a 1.0-L container.

a. Write the equilibrium expression for the reaction.

b. Is the system at equilibrium?

c. If not, will the rate of the forward or reverse reaction initially speed up?

d. At equilibrium, which concentration(s) will be greater than 1.0 mol/L, and which will be less than 1.0 mol/L?

13.80 Consider the following reaction: (13.3, 13.4, 13.5)

$$
\text{PCl}_3(g) \rightleftharpoons \text{PCl}_5(g) + \text{Cl}_2(g)
$$

a. Write the equilibrium expression for the reaction.

b. Initially, 0.60 mol of PCl₃ is placed in a 1.0-L flask. At equilibrium, there is 0.16 mol of PCl₅ in the flask. What are the equilibrium concentrations of PCl₃ and Cl₂?

c. What is the numerical value of the equilibrium constant, Kₑ, for the reaction?

d. If 0.20 mol of Cl₂ is added to the equilibrium mixture, will the concentration of PCl₅ increase or decrease?

13.81 Indicate how each of the following will affect the equilibrium concentration of CO in the following reaction: (13.3, 13.5)

$$
\text{C(s)} + \text{H}_2\text{O}(g) + 31 \text{kcal} \rightleftharpoons \text{CO}(g) + \text{H}_2(g)
$$

a. adding more H₂(g)

b. increasing the temperature
c. increasing the volume of the container
d. decreasing the volume of the container
e. adding a catalyst

13.82 Indicate how each of the following will affect the equilibrium concentration of NH₃ in the following reaction: (13.3, 13.5)

$$
4\text{NH}_3(g) + 5\text{O}_2(g) \rightleftharpoons 4\text{NO}(g) + 6\text{H}_2\text{O}(g) + 906 \text{kJ}
$$

a. adding more O₂(g)

b. increasing the temperature
c. increasing the volume of the container
d. adding more NO(g)
e. removing some H₂O(g)

13.83 Indicate if you would increase or decrease the temperature of the reaction to increase the yield of the products in each of the following: (13.5)

a. COBr₂(g) + Heat \rightleftharpoons CO(g) + Br₂(g)

b. 2NO₂(g) + O₂(g) \rightleftharpoons 4NO(g) + Heat
c. H₂(g) + Br₂(g) + Heat \rightleftharpoons 2HBr(g)
13.84 Indicate if you would increase or decrease the volume of the container to increase the yield of the products in each of the following: (13.5)

a. \( \text{Cl}_2(g) + 2\text{NO}(g) \rightleftharpoons 2\text{NOCl}(g) \)

b. \( \text{N}_2(g) + 2\text{H}_2(g) \rightleftharpoons \text{N}_2\text{H}_4(g) \)

c. \( \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \)

13.85 The antacid milk of magnesia, which contains \( \text{Mg(OH)}_2 \), is used to neutralize excess stomach acid. If the solubility of \( \text{Mg(OH)}_2 \) in water is \( 9.7 \times 10^{-5} \text{ g/L} \), what is the numerical value of \( K_{sp} \) for \( \text{Mg(OH)}_2 \)? (13.6)

13.86 A saturated solution of calcium hydroxide, \( \text{Ca(OH)}_2 \), has a solubility of 0.173 g in 100 mL of water at 20 °C. What is the numerical value of \( K_{sp} \) for \( \text{Ca(OH)}_2 \) at this temperature?

13.37 a. Equilibrium shifts in the direction of the product.

b. Equilibrium shifts in the direction of the product.

c. Equilibrium shifts in the direction of the product.

d. No shift in equilibrium occurs.

e. Equilibrium shifts in the direction of the reactants.

13.39 a. The oxygen concentration is lowered.

b. Equilibrium shifts in the direction of the reactants.

c. \( \text{MgCO}_3(s) \rightleftharpoons \text{Mg}^{2+}(aq) + \text{CO}_3^{2-}(aq); \)

\[ K_{sp} = [\text{Mg}^{2+}][\text{CO}_3^{2-}] \]

d. \( \text{CaF}_2(s) \rightleftharpoons \text{Ca}^{2+}(aq) + 2\text{F}^-(aq); \)

\[ K_{sp} = [\text{Ca}^{2+}][\text{F}^{-}]^2 \]

e. \( \text{Ag}_3\text{PO}_4(s) \rightleftharpoons 3\text{Ag}^+(aq) + \text{PO}_4^{3-}(aq); \)

\[ K_{sp} = [\text{Ag}^+]^3[\text{PO}_4^{3-}] \]

13.43 \( K_{sp} = 1 \times 10^{-10} \)

13.45 \( K_{sp} = 8.8 \times 10^{-12} \)

13.47 \( S = 1 \times 10^{-6} \text{ mol/L} \)

13.49 a. \( \text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \)

b. \( K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}] \)

13.51 a. \( K_{c} = \left[ \frac{5\text{Cl}_2\text{I}}{[\text{CS}_2][\text{CCl}_4]} \right] \)

b. \( K_{c} = \left[ \frac{\text{NOCl}}{[\text{NO}_2]^2[\text{Cl}_2]} \right] \)

c. \( K_{c} = \left[ \frac{[\text{I}_2][\text{CO}_3]^5}{[\text{CO}_3][][\text{I}_2][\text{O}_3]} \right] \)

(As concentration of pure solids and liquids are equal to 1)

13.53 The equilibrium constant for the reaction would have a small value.

13.55 \( T_2 \) is lower than \( T_1 \).

13.57 a. Equilibrium shifts in the direction of the products.

b. Equilibrium shifts in the direction of the products.

c. Equilibrium does not change.

d. Equilibrium shifts in the direction of the products.

e. Mostly reactants

13.59 a. Mostly reactants

b. Mostly products

c. Mostly reactants

d. Mostly products

13.61 a. \( K_{c} = \left[ \frac{[\text{CO}][\text{Cl}_2]}{[\text{COCl}_2]} \right] \)

b. \( K_{c} = 2.5 \)

13.63 \( \text{O}_3 = 3.2 \text{ M} \)

13.65 a. Equilibrium shifts in the direction of the product.

b. Equilibrium shifts in the direction of the reactant.

c. Equilibrium shifts in the direction of the product.

d. Equilibrium shifts in the direction of the reactant.

13.67 a. Equilibrium shifts in the direction of the product.

b. Equilibrium shifts in the direction of the reactant.

c. Equilibrium shifts in the direction of the product.

d. Equilibrium shifts in the direction of the reactant.

13.69 [COCl_2] = 0.44 M
13.71 a. $\text{KF(s)} \rightleftharpoons K^+(aq) + F^-(aq); \quad K_{sp} = [K^+] [F^-]$
b. $\text{Mg(OH)_2(s)} \rightleftharpoons \text{Mg}^{2+}(aq) + 2\text{OH}^-(aq); \quad K_{sp} = [\text{Mg}^{2+})][\text{OH}^-]^2$
c. $\text{Ca}_3(\text{PO}_4)_2(s) \rightleftharpoons 3\text{Ca}^{2+}(aq) + 2\text{PO}_4^{3-}(aq); \quad K_{sp} = [\text{Ca}^{2+})]^3[\text{PO}_4^{3-})]^2$

13.73 $K_{sp} = 5.47 \times 10^{-12}$
13.75 $K_{sp} = 2.0 \times 10^{-13}$
13.77 $S = 1.0 \times 10^{-12} \text{M}$

13.79 a. $K_c = \frac{[\text{NO}]^2[\text{Br}_2]}{[\text{NOBr}]^2}$
b. When the concentrations are placed in the expression, the result is 1.0, which is not equal to $K_c$. The system is not at equilibrium.

c. The rate of the forward reaction will increase.
d. The $[\text{Br}_2]$ and $[\text{NO}]$ will increase and $[\text{NOBr}]$ will decrease.

13.81 a. decrease  b. increase
c. increase  d. decrease
e. no change

13.83 a. increase  b. decrease
c. increase

c. $K_{sp} = 2.0 \times 10^{-11}$
A 30-year-old man has been brought to the emergency room after an automobile accident. The emergency room nurses are tending to the patient, Larry, who is unresponsive. A blood sample is taken then sent to Brianna, a clinical laboratory technician, who begins the process of analyzing the pH, the partial pressures of $O_2$ and $CO_2$, and the concentrations of glucose and electrolytes.

Within minutes, Brianna determines that Larry’s blood pH is 7.30 and the partial pressure of $CO_2$ gas is above the desired level. Blood pH is typically in the range of 7.35 to 7.45, and a value less than 7.35 indicates a state of acidosis. Respiratory acidosis occurs because an increase in the partial pressure of $CO_2$ gas in the bloodstream prevents the biochemical buffers in blood from making a change in the pH.

Brianna recognizes these signs and immediately contacts the emergency room to inform them that Larry’s airway may be blocked. In the emergency room, they provide Larry with an IV containing bicarbonate to increase the blood pH and begin the process of unblocking his airway. Shortly afterward, Larry’s airway is cleared, and his blood pH and partial pressure of $CO_2$ gas return to normal.

CAREER

Clinical Laboratory Technician

Clinical laboratory technicians, also known as medical laboratory technicians, perform a wide variety of tests on body fluids and cells that help in the diagnosis and treatment of patients. These tests range from determining blood concentrations of glucose and cholesterol to determining drug levels in the blood for transplant patients or a patient undergoing treatment. Clinical laboratory technicians also prepare specimens in the detection of cancerous tumors, and type blood samples for transfusions. Clinical laboratory technicians must also interpret and analyze the test results, which are then passed on to the physician.
14.1 Acids and Bases

**LEARNING GOAL** Describe and name acids and bases.

Acids and bases are important substances in health, industry, and the environment. One of the most common characteristics of acids is their sour taste. Lemons and grapefruits taste sour because they contain acids such as citric and ascorbic acid (vitamin C). Vinegar tastes sour because it contains acetic acid. We produce lactic acid in our muscles when we exercise. Acid from bacteria turns milk sour in the production of yogurt and cottage cheese. We have hydrochloric acid in our stomachs that helps us digest food. Sometimes we take antacids, which are bases such as sodium bicarbonate or milk of magnesia, to neutralize the effects of too much stomach acid.

The term *acid* comes from the Latin word *acidus*, which means “sour.” We are familiar with the sour tastes of vinegar and lemons and other common acids in foods.

In 1887, the Swedish chemist Svante Arrhenius was the first to describe acids as substances that produce hydrogen ions (H⁺) when they dissolve in water. Because acids produce ions in water, they are also electrolytes. For example, hydrogen chloride (HCl) dissolves in water to give hydrogen ions, H⁺, and chloride ions, Cl⁻. The hydrogen ions give acids a sour taste, change the blue litmus indicator to red, and corrode some metals.

\[
\text{HCl(g)} \rightarrow \text{HCl(aq)} \rightarrow \text{H}^+(aq) + \text{Cl}^-(aq)
\]

**Naming Acids**

Acids dissolve in water to produce hydrogen ions, along with a negative ion that may be a simple nonmetal anion or a polyatomic ion. When an acid dissolves in water to produce a hydrogen ion and a simple nonmetal anion, the prefix *hydro* is used before the name of the nonmetal, and its *ide* ending is changed to *ic acid*. For example, hydrogen chloride (HCl) dissolves in water to form HCl(aq), which is named hydrochloric acid. An exception is hydrogen cyanide (HCN), which as an acid is named hydrocyanic acid.

When an acid contains oxygen, it dissolves in water to produce a hydrogen ion and an oxygen-containing polyatomic anion. The most common form of an oxygen-containing acid has a name that ends with *ic acid*. The name of its polyatomic anion ends in *ite*. If the acid contains a polyatomic ion with an *ite* ending, its name ends in *ous acid*. 

*These Key Math Skills and Core Chemistry Skills from previous chapters are listed here for your review as you proceed to the new material in this chapter.*
The halogens in Group 7A (17) can form more than two oxygen-containing acids. For chlorine, the common form is chloric acid (HClO₃), which contains the chlorate polyatomic ion (ClO₃⁻). For the acid that contains one more oxygen atom than the common form, the prefix hyp is used; HClO₄ is named perchloric acid. When the polyatomic ion in the acid has one oxygen atom less than the common form, the suffix o is used. Thus, HClO₂ is named chlorous acid; it contains the chlorite ion (ClO₂⁻). The prefix hypo is used for the acid that has two oxygen atoms less than the common form; HClO₃ is named hypochlorous acid. The names of some common acids and their anions are listed in **TABLE 14.1**.

**TABLE 14.1** Names of Common Acids and Their Anions

<table>
<thead>
<tr>
<th>Acid</th>
<th>Name of Acid</th>
<th>Anion</th>
<th>Name of Anion</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>Hydrochloric acid</td>
<td>Cl⁻</td>
<td>Chloride</td>
</tr>
<tr>
<td>HBr</td>
<td>Hydrobromic acid</td>
<td>Br⁻</td>
<td>Bromide</td>
</tr>
<tr>
<td>HI</td>
<td>Hydroiodic acid</td>
<td>I⁻</td>
<td>Iodide</td>
</tr>
<tr>
<td>HCN</td>
<td>Hydrocyanic acid</td>
<td>CN⁻</td>
<td>Cyanide</td>
</tr>
<tr>
<td>HNO₃</td>
<td>Nitric acid</td>
<td>NO₃⁻</td>
<td>Nitrate</td>
</tr>
<tr>
<td>HNO₂</td>
<td>Nitrous acid</td>
<td>NO₂⁻</td>
<td>Nitrite</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>Sulfuric acid</td>
<td>SO₄²⁻</td>
<td>Sulfate</td>
</tr>
<tr>
<td>H₂SO₃</td>
<td>Sulfurous acid</td>
<td>SO₃²⁻</td>
<td>Sulfite</td>
</tr>
<tr>
<td>H₂CO₃</td>
<td>Carbonic acid</td>
<td>CO₃²⁻</td>
<td>Carbonate</td>
</tr>
<tr>
<td>HC₂H₃O₂</td>
<td>Acetic acid</td>
<td>C₂H₃O₂⁻</td>
<td>Acetate</td>
</tr>
<tr>
<td>H₃PO₄</td>
<td>Phosphoric acid</td>
<td>PO₄³⁻</td>
<td>Phosphate</td>
</tr>
<tr>
<td>H₃PO₃</td>
<td>Phosphorous acid</td>
<td>PO₃³⁻</td>
<td>Phosphite</td>
</tr>
<tr>
<td>HClO₂</td>
<td>Chloric acid</td>
<td>ClO₂⁻</td>
<td>Chlorate</td>
</tr>
<tr>
<td>HClO₃</td>
<td>Chlorous acid</td>
<td>ClO₃⁻</td>
<td>Chlorite</td>
</tr>
</tbody>
</table>

**Bases**

You may be familiar with some household bases such as antacids, drain openers, and oven cleaners. According to the Arrhenius theory, bases are ionic compounds that dissociate into cations and hydroxide ions (OH⁻) when they dissolve in water. They are another example of strong electrolytes. For example, sodium hydroxide is an Arrhenius base that dissociates completely in water to give sodium ions (Na⁺) and hydroxide ions (OH⁻).

Most Arrhenius bases are formed from Groups 1A (1) and 2A (2) metals, such as NaOH, KOH, LiOH, and Ca(OH)₂. The hydroxide ions (OH⁻) give Arrhenius bases common characteristics, such as a bitter taste and a slippery feel. A base turns litmus indicator blue and phenolphthalein indicator pink. **TABLE 14.2** compares some characteristics of acids and bases.

**TABLE 14.2** Some Characteristics of Acids and Bases

<table>
<thead>
<tr>
<th>Characteristic</th>
<th>Acids</th>
<th>Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Arrhenius</td>
<td>Produce H⁺</td>
<td>Produce OH⁻</td>
</tr>
<tr>
<td>Electrolytes</td>
<td>Yes</td>
<td>Yes</td>
</tr>
<tr>
<td>Taste</td>
<td>Sour</td>
<td>Bitter, chalky</td>
</tr>
<tr>
<td>Feel</td>
<td>May sting</td>
<td>Soapy, slippery</td>
</tr>
<tr>
<td>Litmus</td>
<td>Red</td>
<td>Blue</td>
</tr>
<tr>
<td>Phenolphthalein</td>
<td>Colorless</td>
<td>Pink</td>
</tr>
<tr>
<td>Neutralization</td>
<td>Neutralize bases</td>
<td>Neutralize acids</td>
</tr>
</tbody>
</table>

An Arrhenius base produces cations and OH⁻ anions in an aqueous solution.
### 14.1 Names and Formulas of Acids and Bases

#### A. Identify each of the following as an acid or a base and give its name:
1. $H_3PO_4$, ingredient in soft drinks
2. $NaOH$, ingredient in oven cleaner

#### B. Write the formula for each of the following:
1. Magnesium hydroxide, ingredient in antacids
2. Hydrobromic acid, used industrially to prepare bromide compounds

### 14.2 Name each of the following acids or bases:

#### a. $HCl$
#### b. $Ca(OH)_2$
#### c. $H_2SO_4$
#### d. KOH
#### e. $HNO_3$
#### f. $HBr$

### 14.3 Name each of the following acids or bases:

#### a. $HCl$
#### b. $Ca(OH)_2$
#### c. $HClO_4$
#### d. $HNO_3$
#### e. $HBr$
#### f. $HBrO_2$

#### a. Rubidium hydroxide
#### b. Hydrofluoric acid
#### c. Phosphoric acid
#### d. Lithium hydroxide
#### e. Ammonium hydroxide
#### f. Periodic acid

### 14.4 Write formulas for each of the following acids and bases:

#### a. Barium hydroxide
#### b. Hydroiodic acid
#### c. Nitric acid
#### d. Strontium hydroxide
#### e. Acetic acid
#### f. Hypochlorous acid

### TRY IT FIRST

#### SOLUTION

#### a. 1. acid, phosphoric acid  
   2. base, sodium hydroxide

#### b. 1. $Mg(OH)_2$  
   2. $HBr$

### STUDY CHECK 14.1

#### a. Identify as an acid or a base and give the name for $H_2CO_3$.
#### b. Write the formula for iron(III) hydroxide.

#### ANSWER

#### a. acid, carbonic acid  
   b. $Fe(OH)_3$

### QUESTIONS AND PROBLEMS

#### 14.1 Acids and Bases

**LEARNING GOAL** Describe and name acids and bases.

#### 14.1 Indicate whether each of the following statements is characteristic of an acid, a base, or both:

a. has a sour taste  
   b. neutralizes bases  
   c. produces $H^+$ ions in water  
   d. is named barium hydroxide  
   e. is an electrolyte

#### 14.2 Indicate whether each of the following statements is characteristic of an acid, a base, or both:

a. neutralizes acids  
   b. produces $OH^-$ ions in water  
   c. has a slippery feel  
   d. conducts an electrical current in solution  
   e. turns litmus red
14.2 Brønsted–Lowry Acids and Bases

**LEARNING GOAL** Identify conjugate acid–base pairs for Brønsted–Lowry acids and bases.

In 1923, J. N. Brønsted in Denmark and T. M. Lowry in Great Britain expanded the definition of acids and bases to include bases that do not contain OH\(^-\) ions. A Brønsted–Lowry acid can donate a hydrogen ion, H\(^+\), and a Brønsted–Lowry base can accept a hydrogen ion.

A Brønsted–Lowry acid is a substance that donates H\(^+\).

A Brønsted–Lowry base is a substance that accepts H\(^+\).

A free hydrogen ion does not actually exist in water. Its attraction to polar water molecules is so strong that the H\(^+\) bonds to a water molecule and forms a hydronium ion, H\(_3\)O\(^+\):

\[
\text{Water} + \text{Hydrogen ion} \rightarrow \left[ \begin{array}{c} \text{H}^- \\ \text{H} \end{array} \right]^+ \text{Hydronium ion}
\]

We can write the formation of a hydrochloric acid solution as a transfer of H\(^+\) from hydrogen chloride to water. By accepting an H\(^+\) in the reaction, water is acting as a base according to the Brønsted–Lowry concept.

\[
\begin{align*}
\text{HCl} & \quad + \\ \text{H}_2\text{O} & \quad \rightarrow \\ \text{H}_3\text{O}^+ & \quad + \\ & \quad \text{Cl}^-
\end{align*}
\]

Acid (H\(^+\) donor) Base (H\(^+\) acceptor) Acidic solution

In another reaction, ammonia (NH\(_3\)) acts as a base by accepting H\(^+\) when it reacts with water. Because the nitrogen atom of NH\(_3\) has a stronger attraction for H\(^+\) than oxygen, water acts as an acid by donating H\(^+\).

\[
\begin{align*}
\text{NH}_3 & \quad + \\ \text{H}_2\text{O} & \quad \leftrightarrow \\ \text{NH}_4^+ & \quad + \\ & \quad \text{OH}^-
\end{align*}
\]

Base (H\(^+\) acceptor) Acid (H\(^+\) donor) Basic solution

**SAMPLE PROBLEM 14.2 Acids and Bases**

In each of the following equations, identify the reactant that is a Brønsted–Lowry acid and the reactant that is a Brønsted–Lowry base:

a. HBr(aq) + H\(_2\)O(l) \(\rightarrow\) H\(_2\)O\(^+\)(aq) + Br\(^-(aq)\)
b. CN\(^-(aq)\) + H\(_2\)O(l) \(\leftrightarrow\) HCN(aq) + OH\(^-(aq)\)

**TRY IT FIRST**

**SOLUTION**

a. HBr, Brønsted–Lowry acid; H\(_2\)O, Brønsted–Lowry base
b. H\(_2\)O, Brønsted–Lowry acid; CN\(^-\), Brønsted–Lowry base
Conjugate Acid–Base Pairs

According to the Brønsted–Lowry theory, a conjugate acid–base pair consists of molecules or ions related by the loss of one H\(^+\) by an acid, and the gain of one H\(^+\) by a base. Every acid–base reaction contains two conjugate acid–base pairs because an H\(^+\) is transferred in both the forward and reverse directions. When an acid such as HF loses one H\(^+\), the conjugate base F\(^-\) is formed. When the base H\(_2\)O gains an H\(^+\), its conjugate acid, H\(_3\)O\(^+\), is formed.

Because the overall reaction of HF is reversible, the conjugate acid H\(_3\)O\(^+\) can donate H\(^+\) to the conjugate base F\(^-\) and re-form the acid HF and the base H\(_2\)O. Using the relationship of loss and gain of one H\(^+\), we can now identify the conjugate acid–base pairs as HF/F\(^-\) along with H\(_3\)O\(^+\)/H\(_2\)O.

In another reaction, ammonia (NH\(_3\)) accepts H\(^+\) from H\(_2\)O to form the conjugate acid NH\(_4\)\(^+\) and conjugate base OH\(^-\). Each of these conjugate acid–base pairs, NH\(_4\)\(^+\)/NH\(_3\) and H\(_2\)O/OH\(^-\), is related by the loss and gain of one H\(^+\).

In these two examples, we see that water can act as an acid when it donates H\(^+\) or as a base when it accepts H\(^+\). Substances that can act as both acids and bases are amphoteric or amphiprotic. For water, the most common amphoteric substance, the acidic or basic behavior depends on the other reactant. Water donates H\(^+\) when it reacts with a stronger base, and it accepts H\(^+\) when it reacts with a stronger acid. Another example of an amphoteric substance is bicarbonate (HCO\(_3\)\(^-\)). With a base, HCO\(_3\)\(^-\) acts as an acid and donates one H\(^+\) to give CO\(_3\)\(^{2-}\). However, when HCO\(_3\)\(^-\) reacts with an acid, it acts as a base and accepts one H\(^+\) to form H\(_2\)CO\(_3\).
SAMPLE PROBLEM 14.3 Identifying Conjugate Acid–Base Pairs

Identify the conjugate acid–base pairs in the following reaction:

\[ \text{HBr}(aq) + \text{NH}_3(aq) \rightarrow \text{Br}^-(aq) + \text{NH}_4^+(aq) \]

**TRY IT FIRST**

**ANSWER**
The conjugate acid–base pairs are HCN/CN⁻ and HSO₄⁻/SO₄²⁻.

STUDY CHECK 14.3

Identify the conjugate acid–base pairs in the following reaction:

\[ \text{HCN}(aq) + \text{SO}_4^{2-}(aq) \leftrightarrow \text{CN}^-(aq) + \text{HSO}_4^-(aq) \]

**ANSWER**
The conjugate acid–base pairs are HCN/CN⁻ and HSO₄⁻/SO₄²⁻.

QUESTIONS AND PROBLEMS

14.2 Brønsted–Lowry Acids and Bases

**LEARNING GOAL** Identify conjugate acid–base pairs for Brønsted–Lowry acids and bases.

14.7 Identify the reactant that is a Brønsted–Lowry acid and the reactant that is a Brønsted–Lowry base in each of the following:
   a. HI(aq) + H₂O(l) \(\rightarrow\) I⁻(aq) + H₃O⁺(aq)
   b. F⁻(aq) + H₂O(l) \(\leftrightarrow\) HF(aq) + OH⁻(aq)
   c. H₂S(aq) + CH₃−CH₂−NH₃(aq) \(\rightarrow\) HS⁻(aq) + CH₃−CH₂−NH₄⁺(aq)
   d. CO₂²⁻(aq) + H₂(l) \(\leftrightarrow\) H₂CO₃⁻(aq) + OH⁻(aq)
   e. H₂SO₄(aq) + H₂O(l) \(\leftrightarrow\) H₂SO₄⁻(aq) + H₂O⁺(aq)
   f. CH₃CO₂⁻(aq) + H₂O(l) \(\leftrightarrow\) CH₃CO₂⁻(aq) + H₂O(l)

14.8 Identify the reactant that is a Brønsted–Lowry acid and the reactant that is a Brønsted–Lowry base in each of the following:
   a. SO₄²⁻(aq) + H⁺(aq) \(\rightarrow\) SO₄³⁻(aq)
   b. CN⁻(aq) + H⁺(aq) \(\rightarrow\) CN⁺(aq) + H₂O(aq)
   c. H⁺(aq) + NO₂⁻(aq) \(\rightarrow\) NO₂⁺(aq) + H₂O(aq)
   d. CO₂⁻(aq) + H⁺(aq) \(\rightarrow\) CO₂⁺(aq) + H₂O(aq)

14.9 Write the formula for the conjugate base for each of the following acids:
   a. HF       b. H₂O       c. H₂PO₃⁻
   d. HSO₄⁻    e. HClO₂

14.10 Write the formula for the conjugate base for each of the following acids:
   a. HCO₃⁻    b. CH₃−NH₃⁺  c. HPO₄²⁻
   d. HNO₂     e. HBrO

14.11 Write the formula for the conjugate acid for each of the following bases:
   a. CO₂²⁻    b. H₂O       c. H₃PO₄⁻
   d. Br⁻       e. ClO₄⁻

14.12 Write the formula for the conjugate acid for each of the following bases:
   a. SO₄³⁻    b. CN⁻    c. NH₃
   d. ClO₂⁻    e. HS⁻

14.13 Identify the Brønsted–Lowry acid–base pairs in each of the following equations:
   a. H⁺(aq) + H₂O(l) \(\rightarrow\) H₂O⁺(aq)
   b. NH₃(aq) + H₂O(l) \(\leftrightarrow\) NH₄⁺(aq) + H₂O(aq)
   c. HCN(aq) + NO₂⁻(aq) \(\rightarrow\) CN⁻(aq) + HNO₂(aq)
   d. CH₂O₂⁻(aq) + H⁺(aq) \(\rightarrow\) CH₂O₂⁻(aq) + H₂O(l)

14.14 Identify the Brønsted–Lowry acid–base pairs in each of the following equations:
   a. H⁺(aq) + H₂O(l) \(\rightarrow\) H₂O⁺(aq)
   b. NH₃(aq) + H⁺(aq) \(\rightarrow\) NH₄⁺(aq) + H₂O(aq)
   c. HCN(aq) + NO₂⁻(aq) \(\rightarrow\) CN⁻(aq) + HNO₂(aq)
   d. HNO₂(aq) + CH₃−CH₂−NH₃(aq) \(\rightarrow\) CH₃−CH₂−NH₄⁺(aq) + NO₂⁻(aq)

14.15 When ammonium chloride dissolves in water, the ammonium ion NH₄⁺ donates an H⁺ to water. Write a balanced equation for the reaction of the ammonium ion with water.

14.16 When sodium carbonate dissolves in water, the carbonate ion CO₃²⁻ acts as a base. Write a balanced equation for the reaction of the carbonate ion with water.
In the process called dissociation, an acid or a base separates into ions in water. The strength of an acid is determined by the moles of \( \text{H}_3\text{O}^+ \) that are produced for each mole of acid that dissolves. The strength of a base is determined by the moles of \( \text{OH}^- \) that are produced for each mole of base that dissolves. Strong acids and strong bases dissociate completely in water, whereas weak acids and weak bases dissociate only slightly, leaving most of the initial acid or base undissociated.

### Strong and Weak Acids

**Strong acids** are examples of strong electrolytes because they donate \( \text{H}^+ \) so easily that their dissociation in water is essentially complete. For example, when HCl, a strong acid, dissociates in water, \( \text{H}^+ \) is transferred to \( \text{H}_2\text{O} \); the resulting solution contains essentially only the ions \( \text{H}_3\text{O}^+ \) and \( \text{Cl}^- \). We consider the reaction of HCl in \( \text{H}_2\text{O} \) as going 100% to products.

Thus, one mole of a strong acid dissociates in water to yield one mole of \( \text{H}_3\text{O}^+ \) and one mole of its conjugate base. We write the equation for a strong acid such as HCl with a single arrow.

\[
\text{HCl}(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)
\]

There are only six common strong acids, which are stronger acids than \( \text{H}_3\text{O}^+ \). All other acids are weak. **Table 14.3** lists the relative strengths of acids and bases. **Weak**

**Table 14.3 Relative Strengths of Acids and Bases**

<table>
<thead>
<tr>
<th>Acid</th>
<th>Conjugate Base</th>
<th>Acid</th>
<th>Conjugate Base</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Strong Acids</strong></td>
<td></td>
<td><strong>Weak Acids</strong></td>
<td></td>
</tr>
<tr>
<td>Hydroiodic acid</td>
<td>( \text{HI} )</td>
<td>Hydrogen sulfate ion</td>
<td>( \text{HSO}_4^- )</td>
</tr>
<tr>
<td>Hydrobromic acid</td>
<td>( \text{HBr} )</td>
<td>Phosphoric acid</td>
<td>( \text{H}_3\text{PO}_4 )</td>
</tr>
<tr>
<td>Perchloric acid</td>
<td>( \text{HClO}_4 )</td>
<td>Nitrous acid</td>
<td>( \text{HNO}_2 )</td>
</tr>
<tr>
<td>Hydrochloric acid</td>
<td>( \text{HCl} )</td>
<td>Hydrofluoric acid</td>
<td>( \text{HF} )</td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>( \text{H}_2\text{SO}_4 )</td>
<td>Acetic acid</td>
<td>( \text{HC}_2\text{H}_2\text{O}_2 )</td>
</tr>
<tr>
<td>Nitric acid</td>
<td>( \text{HNO}_3 )</td>
<td>Carbonic acid</td>
<td>( \text{H}_2\text{CO}_3 )</td>
</tr>
<tr>
<td>Hydronium ion</td>
<td>( \text{H}_2\text{O}^+ )</td>
<td>Hydrosulfuric acid</td>
<td>( \text{H}_2\text{S} )</td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>( \text{NH}_4^+ )</td>
<td>Dihydrogen phosphate ion</td>
<td>( \text{H}_2\text{PO}_4^- )</td>
</tr>
<tr>
<td>Hydrogen sulfate ion</td>
<td>( \text{HSO}_4^- )</td>
<td>Ammonium ion</td>
<td>( \text{NH}_4^+ )</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>( \text{H}_3\text{PO}_4 )</td>
<td>Hydrocyanic acid</td>
<td>( \text{HCN} )</td>
</tr>
<tr>
<td>Nitrous acid</td>
<td>( \text{HNO}_2 )</td>
<td>Bicarbonate ion</td>
<td>( \text{HCO}_3^- )</td>
</tr>
<tr>
<td>Hydrofluoric acid</td>
<td>( \text{HF} )</td>
<td>Methylammonium ion</td>
<td>( \text{CH}_3\text{NH}_3^+ )</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>( \text{HC}_2\text{H}_2\text{O}_2 )</td>
<td>Hydrogen phosphate ion</td>
<td>( \text{HPO}_4^{2-} )</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>( \text{H}_2\text{CO}_3 )</td>
<td>Water</td>
<td>( \text{H}_2\text{O} )</td>
</tr>
<tr>
<td>Hydrosulfuric acid</td>
<td>( \text{H}_2\text{S} )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Dihydrogen phosphate ion</td>
<td>( \text{H}_2\text{PO}_4^- )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>( \text{NH}_4^+ )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrocyanic acid</td>
<td>( \text{HCN} )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Bicarbonate ion</td>
<td>( \text{HCO}_3^- )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Methylammonium ion</td>
<td>( \text{CH}_3\text{NH}_3^+ )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrogen phosphate ion</td>
<td>( \text{HPO}_4^{2-} )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Water</td>
<td>( \text{H}_2\text{O} )</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Acids are weak electrolytes because they dissociate slightly in water, forming only a small amount of $H_3O^+$ ions. A weak acid has a strong conjugate base, which is why the reverse reaction is more prevalent. Even at high concentrations, weak acids produce low concentrations of $H_3O^+$ ions (see Figure 14.1).

Many of the products you use at home contain weak acids. Citric acid is a weak acid found in fruits and fruit juices such as lemons, oranges, and grapefruit. The vinegar used in salad dressings is typically a 5% ($m/v$) acetic acid ($HCO_2H$) solution. In water, a few $HCO_2H$ molecules donate $H^+$ to $H_2O$ to form $H_3O^+$ ions and acetate ions ($C_2H_3O_2^-$). The reverse reaction also takes place, which converts the $H_3O^+$ ions and acetate ions ($C_2H_3O_2^-$) back to reactants. The formation of hydronium ions from vinegar is the reason we notice the sour taste of vinegar. We write the equation for a weak acid in an aqueous solution with a double arrow to indicate that the forward and reverse reactions are at equilibrium.

$$HCO_2H(aq) + H_2O(l) \rightleftharpoons C_2H_3O_2^-(aq) + H_3O^+(aq)$$

**Acetic acid**

**Acetate ion**

**Diprotic Acids**

Some weak acids, such as carbonic acid, are *diprotic acids* that have two $H^+$, which dissociate one at a time. For example, carbonated soft drinks are prepared by dissolving $CO_2$ in water to form carbonic acid ($H_2CO_3$). A weak acid such as $H_2CO_3$ reaches equilibrium between the mostly undissociated $H_2CO_3$ molecules and the ions $H_3O^+$ and $HCO_3^-$. Because $HCO_3^-$ is also a weak acid, a second dissociation can take place to produce another hydronium ion and the carbonate ion ($CO_3^{2-}$).

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

**Carbonic acid**

**Bicarbonate ion**

(hydrogen carbonate)

Because $HCO_3^-$ is also a weak acid, a second dissociation can take place to produce another hydronium ion and the carbonate ion ($CO_3^{2-}$).

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$$

**Bicarbonate ion**

(hydrogen carbonate)

**Carbonate ion**

Carbonic acid, a weak acid, loses one $H^+$ to form hydrogen carbonate ion, which loses a second $H^+$ to form carbonate ion.

**Figure 14.1** A strong acid such as HCl is completely dissociated, whereas a weak acid such as $HCO_2H$ contains mostly molecules and a few ions.

What is the difference between a strong acid and a weak acid?
Sulfuric acid (H$_2$SO$_4$) is also a diprotic acid. However, its first dissociation is complete (100%), which means H$_2$SO$_4$ is a strong acid. The product, hydrogen sulfate (HSO$_4^-$) can dissociate again but only slightly, which means that the hydrogen sulfate ion is a weak acid.

\[
\text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{HSO}_4^-(aq)
\]

**Sulfuric acid**

**Bisulfate ion**

(transported as hydrogen sulfate)

\[
\text{HSO}_4^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{SO}_4^{2-}(aq)
\]

**Bisulfate ion**

(transported as hydrogen sulfate)

**Sulfate ion**

In summary, a strong acid such as HI in water dissociates completely to form an aqueous solution of the ions H$_3$O$^+$ and I$^-$. A weak acid such as HF dissociates only slightly in water to form an aqueous solution that consists mostly of HF molecules and a few H$_3$O$^+$ and F$^-$ ions (see **FIGURE 14.2**).

- **Strong acid:** HI(aq) + H$_2$O(l) $\rightarrow$ H$_3$O$^+(aq)$ + I$^-(aq)$  **Completely dissociated**
- **Weak acid:** HF(aq) + H$_2$O(l) $\rightleftharpoons$ H$_3$O$^+(aq)$ + F$^-(aq)$  **Slightly dissociated**

**FIGURE 14.2** After dissociation in water, **(a)** the strong acid HI has high concentrations of H$_3$O$^+$ and I$^-$, and **(b)** the weak acid HF has a high concentration of HF and low concentrations of H$_3$O$^+$ and F$^-$.  

- **How do the heights of H$_3$O$^+$ and F$^-$ compare to the height of the weak acid HF in the bar diagram for HF?**

**Strong and Weak Bases**

As strong electrolytes, **strong bases** dissociate completely in water. Because these strong bases are ionic compounds, they dissociate in water to give an aqueous solution of metal ions and hydroxide ions. The Group 1A (1) hydroxides are very soluble in water, which can give high concentrations of OH$^-$ ions. A few strong bases are less soluble in water, but what does dissolve dissociates completely as ions. For example, when KOH forms a KOH solution, it contains only the ions K$^+$ and OH$^-$.

\[
\text{KOH}(s) \rightarrow \text{K}^+(aq) + \text{OH}^-(aq)
\]
Strong Bases
Lithium hydroxide (LiOH)
Sodium hydroxide (NaOH)
Potassium hydroxide (KOH)
Rubidium hydroxide (RbOH)
Cesium hydroxide (CsOH)
Calcium hydroxide (Ca(OH)₂)*
Strontium hydroxide (Sr(OH)₂)*
Barium hydroxide (Ba(OH)₂)*

*Low solubility

Weak bases are weak electrolytes that are poor acceptors of hydrogen ions and produce very few ions in solution. A typical weak base, ammonia (NH₃) is found in window cleaners. In an aqueous solution, only a few ammonia molecules accept hydrogen ions to form NH₄⁺ and OH⁻.

\[
\text{NH}_3(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)
\]

Ammonia Ammonium hydroxide

Direction of Reaction
There is a relationship between the components in each conjugate acid–base pair. Strong acids have weak conjugate bases that do not readily accept H⁺. As the strength of the acid decreases, the strength of its conjugate base increases.

In any acid–base reaction, there are two acids and two bases. However, one acid is stronger than the other acid, and one base is stronger than the other base. By comparing their relative strengths, we can determine the direction of the reaction. For example, the strong acid H₂SO₄ readily gives up H⁺ to water. The hydronium ion H₃O⁺ produced is a weaker acid than H₂SO₄, and the conjugate base HSO₄⁻ is a weaker base than water.

\[
\text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{HSO}_4^-(aq)
\]

Almost products

Stronger acid weaker base

Let’s look at another reaction in which water donates one H⁺ to carbonate (CO₃²⁻) to form HCO₃⁻ and OH⁻. From Table 14.3, we see that HCO₃⁻ is a stronger acid than H₂O. We also see that OH⁻ is a stronger base than CO₃²⁻. To reach equilibrium, the stronger acid and stronger base react in the direction of the weaker acid and weaker base.

\[
\text{CO}_3^{2-}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCO}_3^-(aq) + \text{OH}^-(aq)
\]

Mostly reactants

Weaker base stronger acid

SAMPLE PROBLEM 14.4 Direction of Reaction

Does the equilibrium mixture of the following reaction contain mostly reactants or products?

\[
\text{HF}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq)
\]

TRY IT FIRST

SOLUTION

From Table 14.3, we see that HF is a weaker acid than H₂O⁺ and that H₂O is a weaker base than F⁻. Thus, the equilibrium mixture contains mostly reactants.

\[
\text{HF}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq)
\]

Mostly reactants

Weaker acid weaker base

Bases used in household products are used to remove grease and to open drains.

Bases Used in Household Products

Weak Bases
Window cleaner, ammonia, NH₃
Bleach, NaOCl
Laundry detergent, Na₂CO₃, Na₃PO₄
Toothpaste and baking soda, NaHCO₃

Strong Bases

Drain cleaner, oven cleaner, NaOH

Lime for lawns and agriculture, CaCO₃
Laxatives, antacids, Mg(OH)₂, Al(OH)₃
14.3 Strengths of Acids and Bases

LEARNING GOAL Write equations for the dissociation of strong and weak acids; identify the direction of reaction.

14.17 What is meant by the phrase “A strong acid has a weak conjugate base”?
14.18 What is meant by the phrase “A weak acid has a strong conjugate base”?
14.19 Identify the stronger acid in each of the following pairs:
   a. HBr or HNO₂
   b. H₃PO₄ or HSO₄⁻
   c. HCN or H₂CO₃
14.20 Identify the stronger acid in each of the following pairs:
   a. NH₄⁺ or H₂O
   b. H₂SO₄ or HCl
   c. H₂O or H₂CO₃
14.21 Identify the weaker acid in each of the following pairs:
   a. HCl or HSO₄⁻
   b. HNO₂ or HF
   c. HCO₃⁻ or NH₄⁺
14.22 Identify the weaker acid in each of the following pairs:
   a. HNO₃ or HCO₃⁻
   b. HSO₄⁻ or H₂O
   c. H₂SO₄⁻ or H₂CO₃
14.23 Predict whether each of the following reactions contains mostly reactants or products at equilibrium:
   a. H₂CO₃(aq) + H₂O(l) ⇌ HCO₃⁻(aq) + H₂O⁺(aq)
   b. NH₄⁺(aq) + H₂O(l) ⇌ NH₃(aq) + H₂O⁺(aq)
   c. HNO₂(aq) + NH₃(aq) ⇌ NO₂⁻(aq) + NH₄⁺(aq)
14.24 Predict whether each of the following reactions contains mostly reactants or products at equilibrium:
   a. H₃PO₄(aq) + H₂O(l) ⇌ H₂PO₄⁻(aq) + H₂O⁺(aq)
   b. CO₃²⁻(aq) + H₂O(l) ⇌ OH⁻(aq) + HCO₃⁻(aq)
   c. HS⁻(aq) + F⁻(aq) ⇌ HF(aq) + S²⁻(aq)
14.25 Write an equation for the acid–base reaction between ammonium ion and sulfate ion. Why does the equilibrium mixture contain mostly reactants?
14.26 Write an equation for the acid–base reaction between nitrous acid and hydroxide ion. Why does the equilibrium mixture contain mostly products?

14.4 Dissociation Constants for Acids and Bases

LEARNING GOAL Write the dissociation expression for a weak acid or weak base.

As we have seen, acids have different strengths depending on how much they dissociate in water. Because the dissociation of strong acids in water is essentially complete, the reaction is not considered to be an equilibrium situation. However, because weak acids in water dissociate only slightly, the ion products reach equilibrium with the undissociated weak acid molecules. For example, formic acid (HCHO₂) the acid found in bee and ant stings, is a weak acid. Formic acid is a weak acid that dissociates in water to form hydronium ion (H₂O⁺) and formate ion (CHO₂⁻).

\[
HCHO₂(aq) + H₂O(l) \rightleftharpoons H₃O⁺(aq) + CHO₂⁻(aq)
\]

Formic acid, a weak acid, loses one H⁺ to form formate ion.
Writing Dissociation Constants

An acid dissociation expression, $K_a$, can be written for weak acids that gives the ratio of the concentrations of products to the weak acid reactants. As with other dissociation expressions, the molar concentration of the products is divided by the molar concentration of the reactants. Because water is a pure liquid with a constant concentration, it is omitted. The numerical value of the acid dissociation expression is the acid dissociation constant. For example, the acid dissociation expression for the equilibrium equation of formic acid shown above is written

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]}$$

The numerical value of the $K_a$ for formic acid at 25 °C is determined by experiment to be $1.8 \times 10^{-4}$. Thus, for the weak acid HCHO$_2$, the $K_a$ is written

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = 1.8 \times 10^{-4} \quad \text{Acid dissociation constant}$$

The $K_a$ for formic acid is small, which confirms that the equilibrium mixture of formic acid in water contains mostly reactants and only small amounts of the products. (Recall that the brackets in the $K_a$ represent the molar concentrations of the reactants and products). Weak acids have small $K_a$ values. However, strong acids, which are essentially 100% dissociated, have very large $K_a$ values, but these values are not usually given. **TABLE 14.4** gives $K_a$ and $K_b$ values for selected weak acids and bases.

**TABLE 14.4** $K_a$ and $K_b$ Values for Selected Weak Acids and Bases

<table>
<thead>
<tr>
<th>Acids</th>
<th>$K_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Phosphoric acid</td>
<td>$H_3PO_4$</td>
</tr>
<tr>
<td>Nitrous acid</td>
<td>$HNO_2$</td>
</tr>
<tr>
<td>Hydrofluoric acid</td>
<td>$HF$</td>
</tr>
<tr>
<td>Formic acid</td>
<td>$HCHO_2$</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>$HC_2H_3O_2$</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>$H_2CO_3$</td>
</tr>
<tr>
<td>Hydrosulfuric acid</td>
<td>$H_2S$</td>
</tr>
<tr>
<td>Dihydrogen phosphate</td>
<td>$H_2PO_4^-$</td>
</tr>
<tr>
<td>Hydrocyanic acid</td>
<td>$HCN$</td>
</tr>
<tr>
<td>Hydrogen carbonate</td>
<td>$HCO_3^-$</td>
</tr>
<tr>
<td>Hydrogen phosphate</td>
<td>$HPO_4^{2-}$</td>
</tr>
<tr>
<td>Bases</td>
<td>$K_b$</td>
</tr>
<tr>
<td>Methylamine</td>
<td>$CH_3\text{---NH}_2$</td>
</tr>
<tr>
<td>Carbonate</td>
<td>$CO_3^{2-}$</td>
</tr>
<tr>
<td>Ammonia</td>
<td>$NH_3$</td>
</tr>
</tbody>
</table>

Let us now consider the dissociation of the weak base methylamine:

$$CH_3\text{---NH}_2(aq) + H_2O(l) \rightleftharpoons CH_3\text{---NH}_3^+(aq) + OH^- (aq)$$

As we did with the acid dissociation expression, the concentration of water is omitted from the base dissociation expression, $K_b$. The base dissociation constant for methylamine is written

$$K_b = \frac{[CH_3\text{---NH}_3^+][OH^-]}{[CH_3\text{---NH}_2]} = 4.4 \times 10^{-4}$$

**ENGAGE**

Why is an acid with a $K_a = 1.8 \times 10^{-5}$ a stronger acid than an acid with a $K_a = 6.2 \times 10^{-8}$?
TABLE 14.5 summarizes the characteristics of acids and bases in terms of strength and equilibrium position.

TABLE 14.5 Characteristics of Acids and Bases

<table>
<thead>
<tr>
<th>Characteristic</th>
<th>Strong Acids</th>
<th>Weak Acids</th>
</tr>
</thead>
<tbody>
<tr>
<td>Equilibrium Position</td>
<td>Toward products</td>
<td>Toward reactants</td>
</tr>
<tr>
<td>$K_a$</td>
<td>Large</td>
<td>Small</td>
</tr>
<tr>
<td>$[H_3O^+]$ and $[A^-]$</td>
<td>100% of [HA] dissociates</td>
<td>Small percent of [HA] dissociates</td>
</tr>
<tr>
<td>Conjugate Base</td>
<td>Weak</td>
<td>Strong</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Characteristic</th>
<th>Strong Bases</th>
<th>Weak Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Equilibrium Position</td>
<td>Toward products</td>
<td>Toward reactants</td>
</tr>
<tr>
<td>$K_b$</td>
<td>Large</td>
<td>Small</td>
</tr>
<tr>
<td>$[BH^+]$ and $[OH^-]$</td>
<td>100% of [B] reacts</td>
<td>Small percent of [B] reacts</td>
</tr>
<tr>
<td>Conjugate Acid</td>
<td>Weak</td>
<td>Strong</td>
</tr>
</tbody>
</table>

SAMPLE PROBLEM 14.5  Writing an Acid Dissociation Expression

Write the acid dissociation expression for the weak acid nitrous acid.

TRY IT FIRST

SOLUTION

STEP 1 Write the balanced chemical equation. The equation for the dissociation of nitrous acid is written

$$\text{HNO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{NO}_2^-(aq)$$

STEP 2 Write the concentrations of the products as the numerator and the reactants as the denominator. The acid dissociation expression is written as the concentration of the products divided by the concentration of the undissociated weak acid.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$$

STUDY CHECK 14.5

Write the acid dissociation expression for hydrogen phosphate ($\text{HPO}_4^{2-}$).

ANSWER

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{PO}_4^{3-}]}{[\text{HPO}_4^{2-}]}$$

QUESTIONS AND PROBLEMS

14.4 Dissociation Constants for Acids and Bases

LEARNING GOAL Write the dissociation expression for a weak acid or weak base.

14.27 Answer true or false for each of the following: A strong acid  
   a. is completely dissociated in aqueous solution  
   b. has a small value of $K_a$  
   c. has a strong conjugate base  
   d. has a weak conjugate base  
   e. is slightly dissociated in aqueous solution  

14.28 Answer true or false for each of the following: A weak acid  
   a. is completely dissociated in aqueous solution  
   b. has a small value of $K_a$  
   c. has a strong conjugate base  
   d. has a weak conjugate base  
   e. is slightly dissociated in aqueous solution
14.29 Consider the following acids and their dissociation constants:

\[ \text{H}_2\text{SO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HSO}_3^-(aq) \]

\[ K_a = 1.2 \times 10^{-2} \]

\[ \text{HS}^- (aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{O}^+(aq) + \text{S}^{2-} (aq) \]

\[ K_a = 1.3 \times 10^{-19} \]

a. Which is the stronger acid, H$_2$SO$_3$ or HS$^-$?
b. What is the conjugate base of H$_2$SO$_3$?
c. Which acid has the weaker conjugate base?
d. Which acid has the stronger conjugate base?
e. Which acid produces more ions?

14.30 Consider the following acids and their dissociation constants:

\[ \text{HPO}_4^{2-}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{PO}_4^{3-} (aq) \]

\[ K_a = 2.2 \times 10^{-13} \]

\[ \text{HCHO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CHO}_2^- (aq) \]

\[ K_a = 1.8 \times 10^{-4} \]

a. Which is the weaker acid, HPO$_4^{2-}$ or HCHO$_2$?
b. What is the conjugate base of HPO$_4^{2-}$?
c. Which acid has the weaker conjugate base?
d. Which acid has the stronger conjugate base?
e. Which acid produces more ions?

14.31 Phosphoric acid dissociates to form hydronium ion and dihydrogen phosphate. Phosphoric acid has a $K_a$ of $7.5 \times 10^{-3}$.

Write the equation for the reaction and the acid dissociation expression for phosphoric acid.

14.32 Aniline, C$_6$H$_5$—NH$_2$, a weak base with a $K_b$ of $4.0 \times 10^{-10}$, reacts with water to form C$_6$H$_5$—NH$_3^+$ and hydroxide ion.

Write the equation for the reaction and the base dissociation expression for aniline.

14.5 Dissociation of Water

**LEARNING GOAL** Use the water dissociation expression to calculate the [H$_3$O$^+$] and [OH$^-$] in an aqueous solution.

In many acid–base reactions, water is *amphoteric*, which means that it can act either as an acid or as a base. In pure water, there is a forward reaction between two water molecules that transfers H$^+$ from one water molecule to the other. One molecule acts as an acid by losing H$^+$, and the water molecule that gains H$^+$ acts as a base. Every time H$^+$ is transferred between two water molecules, the products are one H$_3$O$^+$ and one OH$^-$, which react in the reverse direction to re-form two water molecules. Thus, equilibrium is reached between the conjugate acid–base pairs of water.

![Conjugate acid–base pair diagram]

**Writing the Water Dissociation Expression, $K_w$**

Using the equation for water at equilibrium, we can write its equilibrium expression that shows the concentrations of the products divided by the concentrations of the reactants. Recall that square brackets around the symbols indicate their concentrations in moles per liter (M).

\[ \text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \]

\[ K = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}][\text{H}_2\text{O}]} \]
By omitting the constant concentration of pure water, we can write the water dissociation expression, \( K_w \):

\[
K_w = [\text{H}_3\text{O}^+][\text{OH}^-]
\]

Experiments have determined, that in pure water, the concentration of \( \text{H}_3\text{O}^+ \) and \( \text{OH}^- \) at 25 °C are each \( 1.0 \times 10^{-7} \) M.

\[
\text{Pure water} \quad [\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}
\]

When we place the \( [\text{H}_3\text{O}^+] \) and \( [\text{OH}^-] \) into the water dissociation expression, we obtain the numerical value of \( K_w \), which is \( 1.0 \times 10^{-14} \) at 25 °C. As before, the concentration units are omitted in the \( K_w \) value.

\[
K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = [1.0 \times 10^{-7}][1.0 \times 10^{-7}] = 1.0 \times 10^{-14}
\]

**Neutral, Acidic, and Basic Solutions**

The \( K_w \) value (\( 1.0 \times 10^{-14} \)) applies to any aqueous solution at 25 °C because all aqueous solutions contain both \( \text{H}_3\text{O}^+ \) and \( \text{OH}^- \) (see **FIGURE 14.3**). When the \( [\text{H}_3\text{O}^+] \) and \( [\text{OH}^-] \) in a solution are equal, the solution is **neutral**. However, most solutions are not neutral; they have different concentrations of \( [\text{H}_3\text{O}^+] \) and \( [\text{OH}^-] \). If acid is added to water, there is an increase in \( [\text{H}_3\text{O}^+] \) and a decrease in \( [\text{OH}^-] \), which makes an acidic solution. If base is added, \( [\text{OH}^-] \) increases and \( [\text{H}_3\text{O}^+] \) decreases, which gives a basic solution. However, for any aqueous solution, whether it is neutral, acidic, or basic, the product \( [\text{H}_3\text{O}^+][\text{OH}^-] \) is equal to \( K_w \) (\( 1.0 \times 10^{-14} \)) at 25 °C (see **TABLE 14.6**).

**FIGURE 14.3**  In a neutral solution, \( [\text{H}_3\text{O}^+] \) and \( [\text{OH}^-] \) are equal. In acidic solutions, the \( [\text{H}_3\text{O}^+] \) is greater than the \( [\text{OH}^-] \). In basic solutions, the \( [\text{OH}^-] \) is greater than the \( [\text{H}_3\text{O}^+] \).

Is a solution that has a \( [\text{H}_3\text{O}^+] \) of \( 1.0 \times 10^{-3} \) M acidic, basic, or neutral?

**TABLE 14.6** Examples of \( [\text{H}_3\text{O}^+] \) and \( [\text{OH}^-] \) in Neutral, Acidic, and Basic Solutions

<table>
<thead>
<tr>
<th>Type of Solution</th>
<th>([\text{H}_3\text{O}^+])</th>
<th>([\text{OH}^-])</th>
<th>(K_w)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Neutral</td>
<td>(1.0 \times 10^{-7}) M</td>
<td>(1.0 \times 10^{-7}) M</td>
<td>(1.0 \times 10^{-14})</td>
</tr>
<tr>
<td>Acidic</td>
<td>(1.0 \times 10^{-2}) M</td>
<td>(1.0 \times 10^{-12}) M</td>
<td>(1.0 \times 10^{-14})</td>
</tr>
<tr>
<td>Acidic</td>
<td>(2.5 \times 10^{-3}) M</td>
<td>(4.0 \times 10^{-10}) M</td>
<td>(1.0 \times 10^{-14})</td>
</tr>
<tr>
<td>Basic</td>
<td>(1.0 \times 10^{-8}) M</td>
<td>(1.0 \times 10^{-6}) M</td>
<td>(1.0 \times 10^{-14})</td>
</tr>
<tr>
<td>Basic</td>
<td>(5.0 \times 10^{-11}) M</td>
<td>(2.0 \times 10^{-4}) M</td>
<td>(1.0 \times 10^{-14})</td>
</tr>
</tbody>
</table>
Using the $K_w$ to Calculate $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in a Solution

If we know the $[\text{H}_3\text{O}^+]$ of a solution, we can use the $K_w$ to calculate the $[\text{OH}^-]$. If we know the $[\text{OH}^-]$ of a solution, we can calculate $[\text{H}_3\text{O}^+]$ from their relationship in the $K_w$, as shown in Sample Problem 14.6.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} \quad [\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]}$$

**SAMPLE PROBLEM 14.6 Calculating the $[\text{H}_3\text{O}^+]$ of a Solution**

A vinegar solution has a $[\text{OH}^-] = 5.0 \times 10^{-12}$ M at 25 °C. What is the $[\text{H}_3\text{O}^+]$ of the vinegar solution? Is the solution acidic, basic, or neutral?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>$[\text{OH}^-]$</td>
<td>$5.0 \times 10^{-12}$ M</td>
<td>$[\text{H}_3\text{O}^+]$</td>
<td>$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$</td>
</tr>
</tbody>
</table>

**STEP 2** Write the $K_w$ for water and solve for the unknown $[\text{H}_3\text{O}^+]$.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

Solve for $[\text{H}_3\text{O}^+]$ by dividing both sides by $[\text{OH}^-]$.

$$\frac{K_w}{[\text{OH}^-]} = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{OH}^-]} = 1.0 \times 10^{-14}$$

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]}$$

**STEP 3** Substitute the known $[\text{OH}^-]$ into the equation and calculate.

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-12}} = 2.0 \times 10^{-3} \text{ M}$$

Because the $[\text{H}_3\text{O}^+]$ of $2.0 \times 10^{-3}$ M is larger than the $[\text{OH}^-]$ of $5.0 \times 10^{-12}$ M, the solution is acidic.

**STUDY CHECK 14.6**

What is the $[\text{H}_3\text{O}^+]$ of an ammonia cleaning solution with $[\text{OH}^-] = 4.0 \times 10^{-4}$ M? Is the solution acidic, basic, or neutral?

**ANSWER**

$[\text{H}_3\text{O}^+] = 2.5 \times 10^{-11}$ M, basic
**14.5 Dissociation of Water**

**LEARNING GOAL** Use the water dissociation expression to calculate the $[\text{H}_2\text{O}^+]$ and $[\text{OH}^-]$ in an aqueous solution.

14.33 Why are the concentrations of $\text{H}_2\text{O}^+$ and $\text{OH}^-$ equal in pure water?

14.34 What is the meaning and value of $K_w$ at 25 °C?

14.35 In an acidic solution, how does the concentration of $\text{H}_2\text{O}^+$ compare to the concentration of $\text{OH}^-$?

14.36 If a base is added to pure water, why does the $[\text{H}_2\text{O}^+]$ decrease?

14.37 Indicate whether each of the following solutions is acidic, basic, or neutral:
   a. $[\text{H}_2\text{O}^+] = 2.0 \times 10^{-5}$ M
   b. $[\text{H}_2\text{O}^+] = 1.4 \times 10^{-9}$ M
   c. $[\text{OH}^-] = 8.0 \times 10^{-3}$ M
   d. $[\text{OH}^-] = 3.5 \times 10^{-10}$ M

14.38 Indicate whether each of the following solutions is acidic, basic, or neutral:
   a. $[\text{H}_2\text{O}^+] = 6.0 \times 10^{-12}$ M
   b. $[\text{H}_2\text{O}^+] = 1.4 \times 10^{-4}$ M
   c. $[\text{OH}^-] = 5.0 \times 10^{-12}$ M
   d. $[\text{OH}^-] = 4.5 \times 10^{-2}$ M

14.39 Calculate the $[\text{H}_2\text{O}^+]$ of each aqueous solution with the following $[\text{OH}^-]$:
   a. coffee, $1.0 \times 10^{-9}$ M
   b. soap, $1.0 \times 10^{-6}$ M
   c. cleanser, $2.0 \times 10^{-3}$ M
   d. lemon juice, $4.0 \times 10^{-13}$ M

14.40 Calculate the $[\text{H}_2\text{O}^+]$ of each aqueous solution with the following $[\text{OH}^-]$:
   a. NaOH, $1.0 \times 10^{-2}$ M
   b. milk of magnesia, $1.0 \times 10^{-3}$ M
   c. aspirin, $1.8 \times 10^{-11}$ M
   d. seawater, $2.5 \times 10^{-6}$ M

**Applications**

14.41 Calculate the $[\text{OH}^-]$ of each aqueous solution with the following $[\text{H}_2\text{O}^+]$:
   a. stomach acid, $4.0 \times 10^{-2}$ M
   b. urine, $5.0 \times 10^{-6}$ M
   c. orange juice, $2.0 \times 10^{-4}$ M
   d. bile, $7.9 \times 10^{-9}$ M

14.42 Calculate the $[\text{OH}^-]$ of each aqueous solution with the following $[\text{H}_2\text{O}^+]$:
   a. baking soda, $1.0 \times 10^{-8}$ M
   b. blood, $4.2 \times 10^{-8}$ M
   c. milk, $5.0 \times 10^{-7}$ M
   d. pancreatic juice, $4.0 \times 10^{-9}$ M

**14.6 The pH Scale**

**LEARNING GOAL** Calculate pH from $[\text{H}_2\text{O}^+]$; given the pH, calculate the $[\text{H}_2\text{O}^+]$ and $[\text{OH}^-]$ of a solution.

In the environment, the acidity, or pH, of rain, water, and soil can have significant effects. When rain becomes too acidic, it can dissolve marble statues and accelerate the corrosion of metals. In lakes and ponds, the acidity of water can affect the ability of plants and fish to survive. The acidity of soil around plants affects their growth. If the soil pH is too acidic or too basic, the roots of the plant cannot take up some nutrients. Most plants thrive in soil with a nearly neutral pH, although certain plants, such as orchids, camellias, and blueberries, require a more acidic soil.

Personnel working in food processing, medicine, agriculture, spa and pool maintenance, soap manufacturing, and wine making measure the $[\text{H}_2\text{O}^+]$ and $[\text{OH}^-]$ of solutions. Although we have expressed $\text{H}_2\text{O}^+$ and $\text{OH}^-$ as molar concentrations, it is more convenient to describe the acidity of solutions using the pH scale. On this scale, a number between 0 and 14 represents the $\text{H}_2\text{O}^+$ concentration for common solutions. A neutral solution has a pH of 7.0 at 25 °C. An acidic solution has a pH less than 7.0; a basic solution has a pH greater than 7.0 (see **FIGURE 14.4**).

<table>
<thead>
<tr>
<th>Acidic solution</th>
<th>pH &lt; 7.0</th>
<th>$[\text{H}_2\text{O}^+] &gt; 1.0 \times 10^{-7}$ M</th>
</tr>
</thead>
<tbody>
<tr>
<td>Neutral solution</td>
<td>pH = 7.0</td>
<td>$[\text{H}_2\text{O}^+] = 1.0 \times 10^{-7}$ M</td>
</tr>
<tr>
<td>Basic solution</td>
<td>pH &gt; 7.0</td>
<td>$[\text{H}_2\text{O}^+] &lt; 1.0 \times 10^{-7}$ M</td>
</tr>
</tbody>
</table>
On the pH scale, values below 7.0 are acidic, a value of 7.0 is neutral, and values above 7.0 are basic.

Is apple juice an acidic, a basic, or a neutral solution?

When we relate acidity and pH, we are using an inverse relationship, which is when one component increases while the other component decreases. When an acid is added to pure water, the $\left[H_3O^+\right]$ (acidity) of the solution increases but its pH decreases. When a base is added to pure water, it becomes more basic, which means its acidity decreases and the pH increases.

In the laboratory, a pH meter is commonly used to determine the pH of a solution. There are also various indicators and pH papers that turn specific colors when placed in solutions of different pH values. The pH is found by comparing the color on the test paper or the color of the solution to a color chart (see FIGURE 14.5).
Consider the pH of the following body fluids:

<table>
<thead>
<tr>
<th>Body Fluid</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>Stomach acid</td>
<td>1.4</td>
</tr>
<tr>
<td>Pancreatic juice</td>
<td>8.4</td>
</tr>
<tr>
<td>Sweat</td>
<td>4.8</td>
</tr>
<tr>
<td>Urine</td>
<td>5.3</td>
</tr>
<tr>
<td>Cerebrospinal fluid</td>
<td>7.3</td>
</tr>
</tbody>
</table>

a. Place the pH values of the body fluids on the list in order of most acidic to most basic.

b. Which body fluid has the highest \([H_3O^+]\)?

TRY IT FIRST

SOLUTION

a. The most acidic body fluid is the one with the lowest pH, and the most basic is the body fluid with the highest pH: stomach acid (1.4), sweat (4.8), urine (5.3), cerebrospinal fluid (7.3), pancreatic juice (8.4).

b. The body fluid with the highest \([H_3O^+]\) would have the lowest pH value, which is stomach acid.

STUDY CHECK 14.7

Which body fluid has the highest \([OH^-]\)?

ANSWER

The body fluid with the highest \([OH^-]\) would have the highest pH value, which is pancreatic juice.

Calculating the pH of Solutions

The pH scale is a logarithmic scale that corresponds to the \([H_3O^+]\) of aqueous solutions. Mathematically, pH is the negative logarithm (base 10) of the \([H_3O^+]\).

\[
pH = -\log[H_3O^+]
\]

Essentially, the negative powers of 10 in the molar concentrations are converted to positive numbers. For example, a lemon juice solution with \([H_3O^+] = 1.0 \times 10^{-2} \text{ M}\) has a pH of 2.00. This can be calculated using the pH equation:

\[
pH = -\log[1.0 \times 10^{-2}]
\]

\[
pH = -(-2.00)
\]

\[
= 2.00
\]
The number of decimal places in the pH value is the same as the number of significant figures in the $[\text{H}_3\text{O}^+]$. The number to the left of the decimal point in the pH value is the power of 10.

$$[\text{H}_3\text{O}^+] = 1.0 \times 10^{-2} \quad \text{pH} = 2.00$$

Two SFs Two SFs

Because pH is a log scale, a change of one pH unit corresponds to a tenfold change in $[\text{H}_3\text{O}^+]$. It is important to note that the pH decreases as the $[\text{H}_3\text{O}^+]$ increases. For example, a solution with a pH of 2.00 has a $[\text{H}_3\text{O}^+]$ that is ten times greater than a solution with a pH of 3.00 and 100 times greater than a solution with a pH of 4.00. The pH of a solution is calculated from the $[\text{H}_3\text{O}^+]$ by using the log key and changing the sign as shown in Sample Problem 14.8.

**SAMPLE PROBLEM 14.8 Calculating pH from $[\text{H}_3\text{O}^+]$**

Aspirin, which is acetylsalicylic acid, was the first nonsteroidal anti-inflammatory drug (NSAID) used to alleviate pain and fever. If a solution of aspirin has a $[\text{H}_3\text{O}^+] = 1.7 \times 10^{-3}$ M, what is the pH of the solution?

Aspirin, acetylsalicylic acid, is a weak acid.

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>$[\text{H}_3\text{O}^+] = 1.7 \times 10^{-3}$ M</td>
<td>pH</td>
<td>pH equation</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Enter the $[\text{H}_3\text{O}^+]$ into the pH equation and calculate.

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log[1.7 \times 10^{-3}]$$

**Calculator Procedure**

1. $1.7 \text{ EE or EXP } +/-$ 3 $\log$ $+/-$ $=\Rightarrow$ or $+/-(\log) 1.7 \text{ EE or EXP } +/-$ 3 $=\Rightarrow 2.768551079$

Be sure to check the instructions for your calculator. Different calculators can have different methods for pH calculation.
**STEP 3** Adjust the number of SFs on the right of the decimal point. In a pH value, the number to the left of the decimal point is an *exact* number derived from the power of 10. Thus, the two SFs in the coefficient determine that there are two SFs after the decimal point in the pH value.

<table>
<thead>
<tr>
<th>Coefficient</th>
<th>Power of ten</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.7</td>
<td>$10^{-3}$ M</td>
</tr>
</tbody>
</table>

Two SFs          | Exact          | pH = $-\log[1.7 \times 10^{-3}] = 2.77$

**STUDY CHECK 14.8**
What is the pH of bleach with $[H_3O^+] = 4.2 \times 10^{-12}$ M?

**ANSWER**
$pH = 11.38$

When we need to calculate the pH from $[OH^-]$, we use the $K_w$ to calculate $[H_3O^+]$, place it in the pH equation, and calculate the pH of the solution as shown in Sample Problem 14.9.

**SAMPLE PROBLEM 14.9 Calculating pH from $[OH^-]$**
What is the pH of an ammonia solution with $[OH^-] = 3.7 \times 10^{-3}$ M?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>$[OH^-] = 3.7 \times 10^{-3}$ M</td>
<td>pH</td>
<td>$K_w = [H_3O^+][OH^-]$, pH equation</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Enter the $[H_3O^+]$ into the pH equation and calculate. Because $[OH^-]$ is given for the ammonia solution, we have to calculate $[H_3O^+]$. Using the water dissociation expression, $K_w$, we divide both sides by $[OH^-]$ to obtain $[H_3O^+]$.

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$\frac{K_w}{[OH^-]} = \frac{[H_3O^+][OH^-]}{[OH^-]} = [H_3O^+] = \frac{1.0 \times 10^{-14}}{[3.7 \times 10^{-3}]} = 2.7 \times 10^{-12}$$ M

Now, we enter the $[H_3O^+]$ into the pH equation.

$$pH = -\log[H_3O^+] = -\log[2.7 \times 10^{-12}]$$

**Calculator Procedure**

$$2.7 \text{ EE or EXP } [+/-] 12 \text{ log } [+/-] =$$

**Calculator Display**

$$11.56863624$$

or

$$+/- \text{ log } 2.7 \text{ EE or EXP } [+/-] 12 =$$
**STEP 3** Adjust the number of SFs on the right of the decimal point.

\[ 2.7 \times 10^{-12} \text{ M} \quad \text{pH} = 11.57 \]

Two SFs to the right of the decimal point

**STUDY CHECK 14.9**

Calculate the pH of a sample of bile that has \([\text{OH}^-] = 1.3 \times 10^{-6} \text{ M}\).

**ANSWER**

\( \text{pH} = 8.11 \)

**pOH**

The pOH scale is similar to the pH scale except that pOH is associated with the \([\text{OH}^-]\) of an aqueous solution.

\[ \text{pOH} = -\log[\text{OH}^-] \]

Solutions with high \([\text{OH}^-]\) have low pOH values; solutions with low \([\text{OH}^-]\) have high pOH values. In any aqueous solution, the sum of the pH and pOH is equal to 14.00, which is the negative logarithm of the \(K_w\).

\[ \text{pH} + \text{pOH} = 14.00 \]

For example, if the pH of a solution is 3.50, the pOH can be calculated as follows:

\[
\begin{align*}
\text{pH} + \text{pOH} & = 14.00 \\
\text{pOH} & = 14.00 - \text{pH} = 14.00 - 3.50 = 10.50
\end{align*}
\]

A comparison of \([\text{H}_3\text{O}^+]\), \([\text{OH}^-]\), and their corresponding pH and pOH values is given in **TABLE 14.7**.

---

**TABLE 14.7 A Comparison of pH and pOH Values at 25 °C, \([\text{H}_3\text{O}^+]\), and \([\text{OH}^-]\)**

<table>
<thead>
<tr>
<th>pH</th>
<th>([\text{H}_3\text{O}^+])</th>
<th>([\text{OH}^-])</th>
<th>pOH</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>(10^0)</td>
<td>(10^{-14})</td>
<td>14</td>
</tr>
<tr>
<td>1</td>
<td>(10^{-1})</td>
<td>(10^{-13})</td>
<td>13</td>
</tr>
<tr>
<td>2</td>
<td>(10^{-2})</td>
<td>(10^{-12})</td>
<td>12</td>
</tr>
<tr>
<td>3</td>
<td>(10^{-3})</td>
<td>(10^{-11})</td>
<td>11</td>
</tr>
<tr>
<td>4</td>
<td>(10^{-4})</td>
<td>(10^{-10})</td>
<td>10</td>
</tr>
<tr>
<td>5</td>
<td>(10^{-5})</td>
<td>(10^{-9})</td>
<td>9</td>
</tr>
<tr>
<td>6</td>
<td>(10^{-6})</td>
<td>(10^{-8})</td>
<td>8</td>
</tr>
<tr>
<td>7</td>
<td>(10^{-7})</td>
<td>(10^{-7})</td>
<td>7</td>
</tr>
<tr>
<td>8</td>
<td>(10^{-8})</td>
<td>(10^{-6})</td>
<td>6</td>
</tr>
<tr>
<td>9</td>
<td>(10^{-9})</td>
<td>(10^{-5})</td>
<td>5</td>
</tr>
<tr>
<td>10</td>
<td>(10^{-10})</td>
<td>(10^{-4})</td>
<td>4</td>
</tr>
<tr>
<td>11</td>
<td>(10^{-11})</td>
<td>(10^{-3})</td>
<td>3</td>
</tr>
<tr>
<td>12</td>
<td>(10^{-12})</td>
<td>(10^{-2})</td>
<td>2</td>
</tr>
<tr>
<td>13</td>
<td>(10^{-13})</td>
<td>(10^{-1})</td>
<td>1</td>
</tr>
<tr>
<td>14</td>
<td>(10^{-14})</td>
<td>(10^0)</td>
<td>0</td>
</tr>
</tbody>
</table>

---

Acids produce the sour taste of the fruits we eat.
Calculating $[\text{H}_3\text{O}^+]$ from pH

If we are given the pH of the solution and asked to determine the $[\text{H}_3\text{O}^+]$, we need to reverse the calculation of pH.

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

For example, if the pH of a solution is 3.0, we can substitute it into this equation. The number of significant figures in $[\text{H}_3\text{O}^+]$ is equal to the number of decimal places in the pH value.

$$[\text{H}_3\text{O}^+] = 10^{-3.0} = 1 \times 10^{-3} \text{ M}$$

For pH values that are not whole numbers, the calculation requires the use of the $10^x$ key, which is usually a 2nd function key. On some calculators, this operation is done using the inverse log equation as shown in Sample Problem 14.10.

**SAMPLE PROBLEM 14.10  Calculating $[\text{H}_3\text{O}^+]$ from pH**

Calculate $[\text{H}_3\text{O}^+]$ for a urine sample, which has a pH of 7.5.

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH = 7.5</td>
<td>$[\text{H}_3\text{O}^+]$</td>
<td>$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Enter the pH value into the inverse log equation and calculate.

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-7.5}$$

**Calculator Procedure**

1. $2^{nd}$ (log) +/− 7.5 $\rightarrow$ or 7.5 $+/−$ $2^{nd}$ (log) $\rightarrow$

**Calculator Display**

$$3.16227766E−8$$

Be sure to check the instructions for your calculator. Different calculators can have different methods for this calculation.

**STEP 3** Adjust the SFs for the coefficient. Because the pH value 7.5 has one digit to the right of the decimal point, the coefficient for $[\text{H}_3\text{O}^+]$ is written with one SF.

$$[\text{H}_3\text{O}^+] = 3 \times 10^{-8} \text{ M}$$

One SF

**STUDY CHECK**

What are the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ of Diet Coke that has a pH of 3.17?

**ANSWER**

$$[\text{H}_3\text{O}^+] = 6.8 \times 10^{-4} \text{ M}, [\text{OH}^-] = 1.5 \times 10^{-11} \text{ M}$$
Gastric acid, which contains HCl, is produced by parietal cells that line the stomach. When the stomach expands with the intake of food, the gastric glands begin to secrete a strongly acidic solution of HCl. In a single day, a person may secrete 2000 mL of gastric juice, which contains hydrochloric acid, mucins, and the enzymes pepsin and lipase.

The HCl in the gastric juice activates a digestive enzyme from the chief cells called pepsinogen to form pepsin, which breaks down proteins in food entering the stomach. The secretion of HCl continues until the stomach has a pH of about 2, which is the optimum for activating the digestive enzymes without ulcerating the stomach lining. In addition, the low pH destroys bacteria that reach the stomach. Normally, large quantities of viscous mucus are secreted within the stomach to protect its lining from acid and enzyme damage. Gastric acid may also form under conditions of stress when the nervous system activates the production of HCl. As the contents of the stomach move into the small intestine, cells produce bicarbonate that neutralizes the gastric acid until the pH is about 5.

**Questions and Problems**

**14.6 The pH Scale**

**Learning Goal** Calculate pH from $[H_3O^+]$; given the pH, calculate the $[H_3O^+]$ and $[OH^-]$ of a solution.

14.43 Why does a neutral solution have a pH of 7.0?

14.44 If you know the $[OH^-]$, how can you determine the pH of a solution?

14.45 State whether each of the following solutions is acidic, basic, or neutral:
   a. blood plasma, pH 7.38
   b. vinegar, pH 2.8
   c. drain cleaner, pH 2.8
   d. coffee, pH 5.52
   e. tomatoes, pH 4.2
   f. chocolate cake, pH 7.6

14.46 State whether each of the following solutions is acidic, basic, or neutral:
   a. soda, pH 3.22
   b. shampoo, pH 8.3
   c. laundry detergent, pH 4.56
   d. rain, pH 5.8
   e. honey, pH 3.9
   f. cheese, pH 4.9

14.47 A solution with a pH of 3 is 10 times more acidic than a solution with pH 4. Explain.

14.48 A solution with a pH of 10 is 100 times more basic than a solution with pH 8. Explain.

14.49 Calculate the pH of each solution given the following:
   a. $[H_3O^+] = 1 \times 10^{-4}$ M
   b. $[H_3O^+] = 3 \times 10^{-9}$ M
   c. $[OH^-] = 1 \times 10^{-7}$ M
   d. $[OH^-] = 2.5 \times 10^{-11}$ M
   e. $[H_3O^+] = 6.7 \times 10^{-8}$ M
   f. $[OH^-] = 8.2 \times 10^{-4}$ M

14.50 Calculate the pOH of each solution given the following:
   a. $[H_3O^+] = 1 \times 10^{-8}$ M
   b. $[H_3O^+] = 5 \times 10^{-6}$ M
   c. $[OH^-] = 1 \times 10^{-2}$ M
   d. $[OH^-] = 8.0 \times 10^{-3}$ M
   e. $[H_3O^+] = 4.7 \times 10^{-2}$ M
   f. $[OH^-] = 3.9 \times 10^{-6}$ M

**Applications**

14.51 Complete the following table:

<table>
<thead>
<tr>
<th>$[H_3O^+]$</th>
<th>$[OH^-]$</th>
<th>pH</th>
<th>pOH</th>
<th>Acidic, Basic, or Neutral?</th>
</tr>
</thead>
<tbody>
<tr>
<td>$1.0 \times 10^{-6}$ M</td>
<td></td>
<td></td>
<td></td>
<td>3.49</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>2.00</td>
</tr>
</tbody>
</table>

14.52 Complete the following table:

<table>
<thead>
<tr>
<th>$[H_3O^+]$</th>
<th>$[OH^-]$</th>
<th>pH</th>
<th>pOH</th>
<th>Acidic, Basic, or Neutral?</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>10.00</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

14.53 A patient with severe metabolic acidosis has a blood plasma pH of 6.92. What is the $[H_3O^+]$ of the blood plasma?

14.54 A patient with respiratory alkalosis has a blood plasma pH of 7.58. What is the $[H_3O^+]$ of the blood plasma?
14.7 Reactions of Acids and Bases

**LEARNING GOAL** Write balanced equations for reactions of acids with metals, carbonates or bicarbonates, and bases.

Typical reactions of acids and bases include the reactions of acids with metals, carbonates or bicarbonates, and bases. For example, when you drop an antacid tablet in water, the bicarbonate ion and citric acid in the tablet react to produce carbon dioxide bubbles, water, and salt. A salt is an ionic compound that does not have \( \text{H}^+ \) as the cation or \( \text{OH}^- \) as the anion.

**Acids and Metals**

Acids react with certain metals to produce hydrogen gas (\( \text{H}_2 \)) and a salt. Active metals include potassium, sodium, calcium, magnesium, aluminum, zinc, iron, and tin. In these single replacement reactions, the metal ion replaces the hydrogen in the acid.

\[
\begin{align*}
\text{Mg}(s) + 2\text{HCl}(aq) & \rightarrow \text{H}_2(g) + \text{MgCl}_2(aq) \\
\text{Zn}(s) + 2\text{HNO}_3(aq) & \rightarrow \text{H}_2(g) + \text{Zn(NO}_3)_2(aq)
\end{align*}
\]

**Acids React with Carbonates or Bicarbonates**

When an acid is added to a carbonate or bicarbonate, the products are carbon dioxide gas, water, and a salt. The acid reacts with \( \text{CO}_3^{2-} \) or \( \text{HCO}_3^- \) to produce carbonic acid (\( \text{H}_2\text{CO}_3 \)), which breaks down rapidly to \( \text{CO}_2 \) and \( \text{H}_2\text{O} \).

\[
\begin{align*}
2\text{HCl}(aq) + \text{Na}_2\text{CO}_3(aq) & \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + 2\text{NaCl}(aq) \\
\text{HBr}(aq) + \text{NaHCO}_3(aq) & \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{NaBr}(aq)
\end{align*}
\]

**Acids and Hydroxides: Neutralization**

**Neutralization** is a reaction between a strong or weak acid with a strong base to produce water and a salt. The \( \text{H}^+ \) of the acid and the \( \text{OH}^- \) of the base combine to form water. The salt is the combination of the cation from the base and the anion from the acid. We can write the following equation for the neutralization reaction between \( \text{HCl} \) and \( \text{NaOH} \):

\[
\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq)
\]

If we write the strong acid \( \text{HCl} \) and the strong base \( \text{NaOH} \) as ions, we see that \( \text{H}^+ \) combines with \( \text{OH}^- \) to form water, leaving the ions \( \text{Na}^+ \) and \( \text{Cl}^- \) in solution.

\[
\text{H}^+(aq) + \text{Cl}^-(aq) + \text{Na}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l) + \text{Na}^+(aq) + \text{Cl}^-(aq)
\]

When we omit the ions that do not change during the reaction (spectator ions), we obtain the net ionic equation.

\[
\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l) \quad \text{Net ionic equation}
\]

**Balancing Neutralization Equations**

In a neutralization reaction, one \( \text{H}^+ \) always reacts with one \( \text{OH}^- \). Therefore, a neutralization equation may need coefficients to balance the \( \text{H}^+ \) from the acid with the \( \text{OH}^- \) from the base as shown in Sample Problem 14.11.
SAMPLE PROBLEM 14.11 Balancing Equations for Acids

Write the balanced equation for the neutralization of HCl(aq) and Ba(OH)₂(s).

TRY IT FIRST

SOLUTION

STEP 1 Write the reactants and products.

\[ \text{HCl}(aq) + \text{Ba(OH)}_2(s) \rightarrow \text{H}_2\text{O}(l) + \text{salt} \]

STEP 2 Balance the H⁺ in the acid with the OH⁻ in the base. Placing a coefficient of 2 in front of the HCl provides 2H⁺ for the 2OH⁻ from Ba(OH)₂.

\[ 2\text{HCl}(aq) + \text{Ba(OH)}_2(s) \rightarrow \text{H}_2\text{O}(l) + \text{salt} \]

STEP 3 Balance the H₂O with the H⁺ and the OH⁻. Use a coefficient of 2 in front of H₂O to balance 2H⁺ and 2OH⁻.

\[ 2\text{HCl}(aq) + \text{Ba(OH)}_2(s) \rightarrow 2\text{H}_2\text{O}(l) + \text{salt} \]

STEP 4 Write the formula of the salt from the remaining ions. Use the ions Ba²⁺ and 2Cl⁻ and write the formula for the salt as BaCl₂.

\[ 2\text{HCl}(aq) + \text{Ba(OH)}_2(s) \rightarrow 2\text{H}_2\text{O}(l) + \text{BaCl}_2(aq) \]

STUDY CHECK 14.11

Write the balanced equation for the neutralization of H₂SO₄(aq) and LiOH(aq).

ANSWER

\[ \text{H}_2\text{SO}_4(aq) + 2\text{LiOH}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{Li}_2\text{SO}_4(aq) \]

CHEMISTRY LINK TO HEALTH

Antacids

Antacids are substances used to neutralize excess stomach acid (HCl). Some antacids are mixtures of aluminum hydroxide and magnesium hydroxide. These hydroxides are not very soluble in water, so the levels of available OH⁻ are not damaging to the intestinal tract. However, aluminum hydroxide has the side effects of producing constipation and binding phosphate in the intestinal tract, which may cause weakness and loss of appetite. Magnesium hydroxide has a laxative effect. These side effects are less likely when a combination of the antacids is used.

\[ \text{Al(OH)}_3(s) + 3\text{HCl}(aq) \rightarrow 3\text{H}_2\text{O}(l) + \text{AlCl}_3(aq) \]

\[ \text{Mg(OH)}_2(s) + 2\text{HCl}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{MgCl}_2(aq) \]

Some antacids use calcium carbonate to neutralize excess stomach acid. About 10% of the calcium is absorbed into the bloodstream, where it elevates the level of serum calcium. Calcium carbonate is not recommended for patients who have peptic ulcers or a tendency to form kidney stones, which typically consist of an insoluble calcium salt.

\[ \text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{CaCl}_2(aq) \]

Still other antacids contain sodium bicarbonate. This type of antacid neutralizes excess gastric acid, increases blood pH, but also elevates sodium levels in the body fluids. It also is not recommended in the treatment of peptic ulcers.

\[ \text{NaHCO}_3(s) + \text{HCl}(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{NaCl}(aq) \]

The neutralizing substances in some antacid preparations are given in TABLE 14.8.

<table>
<thead>
<tr>
<th>Antacid</th>
<th>Base(s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Amphojel</td>
<td>Al(OH)₃</td>
</tr>
<tr>
<td>Milk of magnesia</td>
<td>Mg(OH)₂</td>
</tr>
<tr>
<td>Mylanta, Maalox, Di-Gel, Gelusil, Riopan</td>
<td>Mg(OH)₂, Al(OH)₃</td>
</tr>
<tr>
<td>Bisodol, Rolaids</td>
<td>CaCO₃, Mg(OH)₂</td>
</tr>
<tr>
<td>Titalac, Tums, Pepto-Bismol</td>
<td>CaCO₃</td>
</tr>
<tr>
<td>Alka-Seltzer</td>
<td>NaHCO₃, KHCO₃</td>
</tr>
</tbody>
</table>
14.55 Complete and balance the equation for each of the following reactions:
   a. ZnCO₃(s) + HBr(aq) →
   b. Zn(s) + HCl(aq) →
   c. HCl(aq) + NaHCO₃(s) →
   d. H₂SO₄(aq) + Mg(OH)₂(s) →

14.56 Complete and balance the equation for each of the following reactions:
   a. KHCO₃(s) + HBr(aq) →
   b. Ca(s) + H₂SO₄(aq) →
   c. H₂SO₄(aq) + Ca(OH)₂(s) →
   d. Na₂CO₃(s) + H₂SO₄(aq) →

14.57 Balance each of the following neutralization reactions:
   a. HCl(aq) + Mg(OH)₂(s) → H₂O(l) + MgCl₂(aq)
   b. H₃PO₄(aq) + LiOH(aq) → H₂O(l) + Li₃PO₄(aq)

14.58 Balance each of the following neutralization reactions:
   a. HNO₃(aq) + Ba(OH)₂(s) → H₂O(l) + Ba(NO₃)₂(aq)
   b. H₂SO₄(aq) + Al(OH)₃(s) → H₂O(l) + Al₂(SO₄)₃(aq)

14.59 Write a balanced equation for the neutralization of each of the following:
   a. H₂SO₄(aq) and NaOH(aq)
   b. HCl(aq) and Fe(OH)₃(s)
   c. H₂CO₃(aq) and Mg(OH)₂(s)

14.60 Write a balanced equation for the neutralization of each of the following:
   a. H₃PO₄(aq) and NaOH(aq)
   b. HI(aq) and LiOH(aq)
   c. HNO₃(aq) and Ca(OH)₂(s)

14.8 Acid–Base Titration

LEARNING GOAL Calculate the molarity or volume of an acid or base solution from titration information.

Suppose we need to find the molarity of a solution of HCl, which has an unknown concentration. We can do this by a laboratory procedure called titration in which we neutralize an acid sample with a known amount of base. In a titration, we place a measured volume of the acid in a flask and add a few drops of an indicator, such as phenolphthalein. An indicator is a compound that dramatically changes color when pH of the solution changes. In an acidic solution, phenolphthalein is colorless. Then we fill a buret with NaOH solution of known molarity and carefully add NaOH solution to neutralize the acid in the flask (see FIGURE 14.6). We know that neutralization has taken place when the phenolphthalein in the solution changes from colorless to pink. This is called the neutralization endpoint. From the measured volume of the NaOH solution and its molarity, we calculate the number of moles of NaOH, the moles of acid, and use the measured volume of acid to calculate its concentration.

FIGURE 14.6 The titration of an acid.
A known volume of an acid is placed in a flask with an indicator and titrated with a measured volume of a base solution, such as NaOH, to the neutralization endpoint.

What data is needed to determine the molarity of the acid in the flask?
SAMPLE PROBLEM 14.12 Titration of an Acid

If 16.3 mL of a 0.185 M Sr(OH)\(_2\) solution is used to titrate the HCl in 0.0250 L of gastric juice, what is the molarity of the HCl solution?

\[
\text{Sr(OH)}_2(aq) + 2\text{HCl}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{SrCl}_2(aq)
\]

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities and concentrations.

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0250 L of HCl solution, 16.3 mL of 0.185 M Sr(OH)(_2) solution</td>
<td>molarity of the HCl solution</td>
<td>molarity, mole–mole factor</td>
</tr>
</tbody>
</table>

**Neutralization Equation**

\[
\text{Sr(OH)}_2(aq) + 2\text{HCl}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{SrCl}_2(aq)
\]

STEP 2 Write a plan to calculate the molarity.

<table>
<thead>
<tr>
<th>mL of Sr(OH)(_2) solution</th>
<th>Metric factor</th>
<th>L of Sr(OH)(_2) solution</th>
<th>Molarity</th>
<th>moles of Sr(OH)(_2)</th>
<th>Mole–mole factor</th>
<th>moles of HCl</th>
<th>Divide by liters</th>
<th>molarity of HCl solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>1000 mL Sr(OH)(_2) solution</td>
<td>1 L Sr(OH)(_2) solution</td>
<td>1000 mL of Sr(OH)(_2) solution</td>
<td>1 L Sr(OH)(_2) solution</td>
<td>0.185 mol Sr(OH)(_2)</td>
<td>0.185 mol Sr(OH)(_2)</td>
<td>1 L Sr(OH)(_2) solution</td>
<td>0.185 mol Sr(OH)(_2)</td>
<td></td>
</tr>
</tbody>
</table>

\[
\frac{2 \text{ mol of HCl}}{1 \text{ mol Sr(OH)}_2} = \frac{1 \text{ mol of Sr(OH)}_2}{2 \text{ mol HCl}}
\]

STEP 3 State equalities and conversion factors, including concentrations.

\[
16.3 \text{ mL Sr(OH)}_2\text{ solution} \times \frac{1 \text{ L Sr(OH)}_2\text{ solution}}{1000 \text{ mL Sr(OH)}_2\text{ solution}} \times \frac{0.185 \text{ mol Sr(OH)}_2}{1 \text{ L Sr(OH)}_2\text{ solution}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Sr(OH)}_2} = 0.00603 \text{ mol of HCl}
\]

\[
\text{molarity of HCl solution} = \frac{0.00603 \text{ mol HCl}}{0.0250 \text{ L HCl solution}} = 0.241 \text{ M HCl solution}
\]

STUDY CHECK 14.12

What is the molarity of an HCl solution if 28.6 mL of a 0.175 M NaOH solution is needed to titrate a 25.0-mL sample of the HCl solution?

**Answer**

0.200 M HCl solution
LEARNING GOAL Calculate the molarity or volume of an acid or base solution from titration information.

14.61 If you need to determine the molarity of a formic acid solution, HCHO₂, how would you proceed?

HCHO₂(aq) + H₂O(l) ⇄ H₃O⁺(aq) + CHO₂⁻(aq)

14.62 If you need to determine the molarity of an acetic acid solution, HC₂H₃O₂, how would you proceed?

HC₂H₃O₂(aq) + H₂O(l) ⇄ H₃O⁺(aq) + C₂H₃O₂⁻(aq)

14.63 What is the molarity of a solution of HCl if 5.00 mL of the HCl solution is titrated with 28.6 mL of a 0.145 M NaOH solution?

HCl(aq) + NaOH(aq) → H₂O(l) + NaCl(aq)

14.64 What is the molarity of an acetic acid solution if 25.0 mL of the HC₂H₃O₂ solution is titrated with 29.7 mL of a 0.205 M KOH solution?

HC₂H₃O₂(aq) + KOH(aq) → H₂O(l) + KC₂H₃O₂(aq)

14.65 If 38.2 mL of a 0.163 M KOH solution is required to neutralize completely 25.0 mL of a solution of H₂SO₄, what is the molarity of the H₂SO₄ solution?

H₂SO₄(aq) + 2KOH(aq) → 2H₂O(l) + K₂SO₄(aq)

14.66 A solution of 0.162 M NaOH is used to titrate 25.0 mL of a solution of H₂SO₄. If 32.8 mL of the NaOH solution is required to reach the endpoint, what is the molarity of the H₂SO₄ solution?

H₂SO₄(aq) + 2NaOH(aq) → 2H₂O(l) + Na₂SO₄(aq)

14.67 A solution of 0.204 M NaOH is used to titrate 50.0 mL of a 0.0224 M H₃PO₄ solution. What volume, in milliliters, of the NaOH solution is required?

H₃PO₄(aq) + 3NaOH(aq) → 3H₂O(l) + Na₃PO₄(aq)

14.68 A solution of 0.312 M KOH is used to titrate 15.0 mL of a 0.186 M H₂PO₄ solution. What volume, in milliliters, of the KOH solution is required?

H₂PO₄(aq) + 3KOH(aq) → 3H₂O(l) + K₃PO₄(aq)

14.69 Buffers

LEARNING GOAL Describe the role of buffers in maintaining the pH of a solution; calculate the pH of a buffer.

The lungs and the kidneys are the primary organs that regulate the pH of body fluids, including blood and urine. Major changes in the pH of the body fluids can severely affect biological activities within the cells. Buffers are present to prevent large fluctuations in pH.

The pH of water and most solutions changes drastically when a small amount of acid or base is added. However, when an acid or a base is added to a buffer solution, there is little change in pH. A buffer solution maintains the pH of a solution by neutralizing small amounts of added acid or base. In the human body, whole blood contains plasma, white blood cells and platelets, and red blood cells. Blood plasma contains buffers that maintain a consistent pH of about 7.4. If the pH of the blood plasma goes slightly above or below 7.4, changes in our oxygen levels and our metabolic processes can be drastic enough to cause death. Even though we obtain acids and bases from foods and cellular reactions, the buffers in the body absorb those compounds so effectively that the pH of our blood plasma remains essentially unchanged (see FIGURE 14.7).

In a buffer, an acid must be present to react with any OH⁻ that is added, and a base must be available to react with any added H₃O⁺. However, that acid and base must not neutralize each other. Therefore, a combination of an acid–base conjugate pair is used in buffers. Most buffer solutions consist of nearly equal concentrations of a weak acid and a salt containing its conjugate base. Buffers may also contain a weak base and the salt of the weak base, which contains its conjugate acid.

For example, a typical buffer can be made from the weak acid acetic acid (HC₂H₃O₂) and its salt, sodium acetate (NaC₂H₃O₂). As a weak acid, acetic acid dissociates slightly in water to form H₃O⁺ and a very small amount of C₂H₃O₂⁻. The addition of its salt, sodium acetate, provides a much larger concentration of acetate ion (C₂H₃O₂⁻), which is necessary for its buffering capability.

HC₂H₃O₂(aq) + H₂O(l) ⇄ H₃O⁺(aq) + C₂H₃O₂⁻(aq)

Large amount Large amount
The buffer described here consists of about equal concentrations of acetic acid (HC$_2$H$_3$O$_2$) and its conjugate base acetate ion (C$_2$H$_3$O$_2^-$). Adding H$_3$O$^+$ to the buffer neutralizes some C$_2$H$_3$O$_2^-$, whereas adding OH$^-$ neutralizes some HC$_2$H$_3$O$_2$. The pH of the solution is maintained as long as the added amount of acid or base is small compared to the concentrations of the buffer components.

How does this acetic acid–acetate ion buffer maintain pH?
ENGAGE
Which part of a buffer neutralizes any $\text{H}_3\text{O}^+$ that is added?

CORE CHEMISTRY SKILL
Calculating the pH of a Buffer

Guide to Calculating pH of a Buffer

STEP 1
State the given and needed quantities.

STEP 2
Write the $K_a$ expression and rearrange for $[\text{H}_3\text{O}^+]$.

STEP 3
Substitute [HA] and [A⁻] into the $K_a$ expression.

STEP 4
Use $[\text{H}_3\text{O}^+]$ to calculate pH.

Calculating the pH of a Buffer

By rearranging the $K_a$ expression to give $[\text{H}_3\text{O}^+]$, we can obtain the ratio of the acetic acid/acetate buffer.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

Solving for $[\text{H}_3\text{O}^+]$ gives:

$$[\text{H}_3\text{O}^+] = K_a \times \frac{[\text{HC}_2\text{H}_3\text{O}_2]}{[\text{C}_2\text{H}_3\text{O}_2^-]}$$

In this rearrangement of $K_a$, the weak acid is in the numerator and the conjugate base in the denominator. We can now calculate the $[\text{H}_3\text{O}^+]$ and pH for an acetic acid buffer as shown in Sample Problem 14.13.

SAMPLE PROBLEM 14.13 Calculating the pH of a Buffer

The $K_a$ for acetic acid, HC$_2$H$_3$O$_2$, is $1.8 \times 10^{-5}$. What is the pH of a buffer prepared with 1.0 M HC$_2$H$_3$O$_2$ and 1.0 M C$_2$H$_3$O$_2^-$?

HC$_2$H$_3$O$_2$(aq) + H$_2$O(l) $\rightleftharpoons$ H$_3$O$^+(aq) +$ C$_2$H$_3$O$_2^-(aq)$

TRY IT FIRST

SOLUTION

STEP 1 State the given and needed quantities.

<table>
<thead>
<tr>
<th>GIVEN</th>
<th>NEED</th>
<th>CONNECT</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.0 M HC$_2$H$_3$O$_2$, 1.0 M C$_2$H$_3$O$_2^-$</td>
<td>pH</td>
<td>$K_a$ expression</td>
</tr>
</tbody>
</table>

Equation

HC$_2$H$_3$O$_2$(aq) + H$_2$O(l) $\rightleftharpoons$ H$_3$O$^+(aq) +$ C$_2$H$_3$O$_2^-(aq)$

STEP 2 Write the $K_a$ expression and rearrange for $[\text{H}_3\text{O}^+]$.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

$$[\text{H}_3\text{O}^+] = K_a \times \frac{[\text{HC}_2\text{H}_3\text{O}_2]}{[\text{C}_2\text{H}_3\text{O}_2^-]}$$

STEP 3 Substitute [HA] and [A⁻] into the $K_a$ expression.

$$[\text{H}_3\text{O}^+] = 1.8 \times 10^{-5} \times \frac{[1.0]}{[1.0]}$$

$$[\text{H}_3\text{O}^+] = 1.8 \times 10^{-5} \text{ M}$$
**STEP 4** Use \([\text{H}_3\text{O}^+]\) to calculate pH. Placing the \([\text{H}_3\text{O}^+]\) into the pH equation gives the pH of the buffer.

\[
\text{pH} = -\log[1.8 \times 10^{-5}] = 4.74
\]

**STUDY CHECK 14.13**

One of the conjugate acid–base pairs that buffers the blood plasma is \(\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}\), which has a \(K_a\) of \(6.2 \times 10^{-8}\). What is the pH of a buffer that is prepared from 0.10 M \(\text{H}_2\text{PO}_4^-\) and 0.50 M \(\text{HPO}_4^{2-}\)?

**ANSWER**

\(\text{pH} = 7.91\)

Because \(K_a\) is a constant at a given temperature, the \([\text{H}_3\text{O}^+]\) is determined by the \([\text{HC}_2\text{H}_3\text{O}_2]/[\text{C}_2\text{H}_3\text{O}_2^-]\) ratio. As long as the addition of small amounts of either acid or base changes the ratio of \([\text{HC}_2\text{H}_3\text{O}_2]/[\text{C}_2\text{H}_3\text{O}_2^-]\) only slightly, the changes in \([\text{H}_3\text{O}^+]\) will be small and the pH will be maintained. If a large amount of acid or base is added, the buffering capacity of the system may be exceeded. Buffers can be prepared from conjugate acid–base pairs such as \(\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}\), \(\text{HPO}_4^{2-}/\text{PO}_4^{3-}\), \(\text{HCO}_3^-/\text{CO}_2^-\), or \(\text{NH}_4^+/:\text{NH}_3\). The pH of the buffer solution will depend on the conjugate acid–base pair chosen.

Using a common phosphate buffer for biological specimens, we can look at the effect of using different ratios of \([\text{H}_2\text{PO}_4^-]/[\text{HPO}_4^{2-}\) on the \([\text{H}_3\text{O}^+]\) and pH. The \(K_a\) of \(\text{H}_2\text{PO}_4^-\) is \(6.2 \times 10^{-8}\). The equation and the \([\text{H}_3\text{O}^+]\) are written as follows:

\[
\text{H}_2\text{PO}_4^- (aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{HPO}_4^{2-} (aq)
\]

\[
[\text{H}_3\text{O}^+] = K_a \times \frac{[\text{H}_2\text{PO}_4^-]}{[\text{HPO}_4^{2-}]}
\]


<table>
<thead>
<tr>
<th>(K_a)</th>
<th>([\text{H}_2\text{PO}_4^-]/[\text{HPO}_4^{2-}])</th>
<th>Ratio</th>
<th>([\text{H}_3\text{O}^+])</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>(6.2 \times 10^{-8})</td>
<td>1.0 M</td>
<td>10</td>
<td>(6.2 \times 10^{-7})</td>
<td>6.21</td>
</tr>
<tr>
<td>(6.2 \times 10^{-8})</td>
<td>0.10 M</td>
<td>1</td>
<td>(6.2 \times 10^{-8})</td>
<td>7.21</td>
</tr>
<tr>
<td>(6.2 \times 10^{-8})</td>
<td>1.0 M</td>
<td>(\frac{1}{10})</td>
<td>(6.2 \times 10^{-9})</td>
<td>8.21</td>
</tr>
</tbody>
</table>

To prepare a phosphate buffer with a pH close to the pH of a biological sample, 7.4, we would choose concentrations that are about equal, such as 1.0 M \(\text{H}_2\text{PO}_4^-\) and 1.0 M \(\text{HPO}_4^{2-}\).

**CHEMISTRY LINK TO HEALTH**

**Buffers in the Blood Plasma**

The arterial blood plasma has a normal pH of 7.35 to 7.45. If changes in \([\text{H}_3\text{O}^+]\) lower the pH below 6.8 or raise it above 8.0, cells cannot function properly and death may result. In our cells, \(\text{CO}_2\) is continually produced as an end product of cellular metabolism. Some \(\text{CO}_2\) is carried to the lungs for elimination, and the rest dissolves in body fluids such as plasma and saliva, forming carbonic acid, \(\text{H}_2\text{CO}_3\). As a weak acid, carbonic acid dissociates to give bicarbonate, \(\text{HCO}_3^-\), and \([\text{H}_3\text{O}^+]\). More of the anion \(\text{HCO}_3^-\) is supplied by the kidneys to give an important buffer system in the body fluid—the \(\text{H}_2\text{CO}_3/\text{HCO}_3^-\) buffer.

\[
\text{CO}_2(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3(aq) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HCO}_3^-(aq)
\]

Excess \([\text{H}_3\text{O}^+]\) entering the body fluids reacts with the \(\text{HCO}_3^-\), and excess \([\text{OH}^-]\) reacts with the carbonic acid.
H₂CO₃(aq) + H₂O(l) ⇌ H₃O⁺(aq) + HCO₃⁻(aq)

Equilibrium shifts in the direction of the reactants

H₂CO₃(aq) + OH⁻(aq) → H₂O(l) + HCO₃⁻(aq)

Equilibrium shifts in the direction of the products

For carbonic acid, we can write the equilibrium expression as

\[ K_a = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} \]

To maintain the normal blood plasma pH (7.35 to 7.45), the ratio of \([H_2CO_3]/[HCO_3^-]\) needs to be about 1 to 10, which is obtained by the concentrations in the blood plasma of 0.0024 M H₂CO₃ and 0.024 M HCO₃⁻.

\[
[H_3O^+] = K_a \times \frac{[H_2CO_3]}{[HCO_3^-]}
\]

\[
[H_3O^+] = 4.3 \times 10^{-7} \times \frac{0.0024}{0.024} = 4.3 \times 10^{-8} M
\]

\[
pH = -\log(4.3 \times 10^{-8}) = 7.37
\]

In the body, the concentration of carbonic acid is closely associated with the partial pressure of CO₂, \(P_{CO_2}\). **TABLE 14.9** lists the normal values for arterial blood. If the CO₂ level rises, increasing \([H_2CO_3]\), the equilibrium shifts to produce more \(H_3O^+\), which lowers the pH. This condition is called *acidosis*. Difficulty with ventilation or gas diffusion can lead to respiratory acidosis, which can happen in emphysema or when an accident or depressive drugs affect the medulla of the brain.

A lowering of the CO₂ level leads to a high blood pH, a condition called *alkalosis*. Excitement, trauma, or a high temperature may cause a person to hyperventilate, which expels large amounts of CO₂. As the partial pressure of CO₂ in the blood falls below normal, the equilibrium shifts from \(H_2CO_3\) to \(CO_2\) and \(H_2O\). This shift decreases the \([H_3O^+]\) and raises the pH. The kidneys also regulate \(H_3O^+\) and \(HCO_3^-\), but they do so more slowly than the adjustment made by the lungs during ventilation.

**TABLE 14.10** lists some of the conditions that lead to changes in the blood pH and some possible treatments.

<table>
<thead>
<tr>
<th>TABLE 14.10 Acidity and Alkalinity: Symptoms, Causes, and Treatments</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Respiratory Acidosis: CO₂↑ pH↓</strong></td>
</tr>
<tr>
<td><strong>Symptoms:</strong> Failure to ventilate, suppression of breathing, disorientation, weakness, coma</td>
</tr>
<tr>
<td><strong>Causes:</strong> Lung disease blocking gas diffusion (e.g., emphysema, pneumonia, bronchitis, asthma); depression of respiratory center by drugs, cardiopulmonary arrest, stroke, poliomyelitis, or nervous system disorders</td>
</tr>
<tr>
<td><strong>Treatment:</strong> Correction of disorder, infusion of bicarbonate</td>
</tr>
<tr>
<td><strong>Metabolic Acidosis: H⁺↑ pH↓</strong></td>
</tr>
<tr>
<td><strong>Symptoms:</strong> Increased ventilation, fatigue, confusion</td>
</tr>
<tr>
<td><strong>Causes:</strong> Renal disease, including hepatitis and cirrhosis; increased acid production in diabetes mellitus, hyperthyroidism, alcoholism, and starvation; loss of alkali in diarrhea; acid retention in renal failure</td>
</tr>
<tr>
<td><strong>Treatment:</strong> Sodium bicarbonate given orally, dialysis for renal failure, insulin treatment for diabetic ketosis</td>
</tr>
<tr>
<td><strong>Respiratory Alkalosis: CO₂↓ pH↑</strong></td>
</tr>
<tr>
<td><strong>Symptoms:</strong> Increased rate and depth of breathing, numbness, light-headedness, tetany</td>
</tr>
<tr>
<td><strong>Causes:</strong> Hyperventilation because of anxiety, hysteria, fever, exercise; reaction to drugs such as salicylate, quinine, and antihistamines; conditions causing hypoxia (e.g., pneumonia, pulmonary edema, heart disease)</td>
</tr>
<tr>
<td><strong>Treatment:</strong> Elimination of anxiety-producing state, rebreathing into a paper bag</td>
</tr>
<tr>
<td><strong>Metabolic Alkalosis: H⁺↓ pH↑</strong></td>
</tr>
<tr>
<td><strong>Symptoms:</strong> Depressed breathing, apathy, confusion</td>
</tr>
<tr>
<td><strong>Causes:</strong> Vomiting, diseases of the adrenal glands, ingestion of excess alkali</td>
</tr>
<tr>
<td><strong>Treatment:</strong> Infusion of saline solution, treatment of underlying diseases</td>
</tr>
</tbody>
</table>
14.9 Buffers

LEARNING GOAL Describe the role of buffers in maintaining the pH of a solution; calculate the pH of a buffer.

14.69 Which of the following represents a buffer system? Explain.
   a. NaOH and NaCl
   b. H₃CO₃ and NaHCO₃
   c. HF and KF
   d. KCl and NaCl

14.70 Which of the following represents a buffer system? Explain.
   a. NaOH and NaCl
   b. H₂CO₃ and NaHCO₃
   c. HF and KF
   d. KCl and NaCl

14.71 Consider the buffer system of hydrofluoric acid, HF, and its salt, NaF.

HF(aq) + H₂O(l) ⇌ H₃O⁺(aq) + F⁻(aq)

a. The purpose of this buffer system is to:
   1. maintain [HF]
   2. maintain [F⁻]
   3. maintain pH
b. The salt of the weak acid is needed to:
   1. provide the conjugate base
   2. neutralize added H₃O⁺
   3. provide the conjugate acid
c. If OH⁻ is added, it is neutralized by:
   1. the salt
   2. H₂O
   3. H₃O⁺
d. When H₃O⁺ is added, the equilibrium shifts in the direction of:
   1. reactants
   2. products
   3. does not change

14.72 Consider the buffer system of nitrous acid, HNO₂, and its salt, NaNO₂.

HNO₂(aq) + H₂O(l) ⇌ H₃O⁺(aq) + NO₂⁻(aq)

a. The purpose of this buffer system is to:
   1. maintain [HNO₂]
   2. maintain [NO₂⁻]
   3. maintain pH
b. The weak acid is needed to:
   1. provide the conjugate base
   2. neutralize added OH⁻
   3. provide the conjugate acid
c. If H₂O⁺ is added, it is neutralized by:
   1. the salt
   2. H₂O
   3. OH⁻
d. When OH⁻ is added, the equilibrium shifts in the direction of:
   1. reactants
   2. products
   3. does not change

14.73 Nitrous acid has a Kₐ of 4.5 × 10⁻⁴. What is the pH of a buffer solution containing 0.10 M HNO₂ and 0.10 M NO₂⁻?

14.74 Acetic acid has a Kₐ of 1.8 × 10⁻⁵. What is the pH of a buffer solution containing 0.15 M HC₂H₃O₂ and 0.15 M C₂H₃O₂⁻?

14.75 Using Table 14.4 for Kₐ values, compare the pH of a HF buffer that contains 0.10 M HF and 0.10 M NaF with another HF buffer that contains 0.060 M HF and 0.120 M NaF.

14.76 Using Table 14.4 for Kₐ values, compare the pH of a H₂CO₃ buffer that contains 0.10 M H₂CO₃ and 0.10 M NaHCO₃ with another H₂CO₃ buffer that contains 0.15 M H₂CO₃ and 0.050 M NaHCO₃.

Applications

14.77 Why would the pH of your blood plasma increase if you breathe fast?

14.78 Why would the pH of your blood plasma decrease if you hold your breath?

14.79 Someone with kidney failure excretes urine with large amounts of HCO₃⁻. How would this loss of HCO₃⁻ affect the pH of the blood plasma?

14.80 Someone with severe diabetes obtains energy by the breakdown of fats, which produce large amounts of acidic substances. How would this affect the pH of the blood plasma?
Larry has not been feeling well lately. He tells his doctor that he has discomfort and a burning feeling in his chest, and a sour taste in his throat and mouth. At times, Larry says he feels bloated after a big meal, has a dry cough, is hoarse, and sometimes has a sore throat. He has tried antacids, but they do not bring any relief.

The doctor tells Larry that he thinks he has acid reflux. At the top of the stomach there is a valve, the lower esophageal sphincter, that normally closes after food passes through it. However, if the valve does not close completely, acid produced in the stomach to digest food can move up into the esophagus, a condition called acid reflux. The acid, which is hydrochloric acid, HCl, is produced in the stomach to kill bacteria, microorganisms, and to activate the enzymes we need to break down food.

If acid reflux occurs, the strong acid HCl comes in contact with the lining of the esophagus, where it causes irritation and produces a burning feeling in the chest. Sometimes the pain in the chest is called heartburn. If the HCl reflux goes high enough to reach the throat, a sour taste may be noticed in the mouth. If Larry's symptoms occur three or more times a week, he may have a chronic condition known as acid reflux disease or gastroesophageal reflux disease (GERD).

Larry's doctor orders an esophageal pH test in which the amount of acid entering the esophagus from the stomach is measured over 24 h. A probe that measures the pH is inserted into the lower esophagus above the esophageal sphincter. The pH measurements indicate a reflux episode each time the pH drops to 4 or less.

In the 24-h period, Larry has several reflux episodes and his doctor determines that he has chronic GERD. He and Larry discuss treatment for GERD, which includes eating smaller meals, not lying down for 3 h after eating, making dietary changes, and losing weight. Antacids may be used to neutralize the acid coming up from the stomach. Other medications known as proton pump inhibitors (PPIs), such as Prilosec and Nexium, may be used to suppress the production of HCl in the stomach (gastric parietal cells), which raises the pH in the stomach to between 4 and 5, and gives the esophagus time to heal. Nexium may be given in oral doses of 40 mg once a day for 4 weeks. In severe GERD cases, an artificial valve may be created at the top of the stomach to strengthen the lower esophageal sphincter.

Applications

14.81 At rest, the \([H_3O^+]\) of the stomach fluid is \(2.0 \times 10^{-4}\) M. What is the pH of the stomach fluid?

14.82 When food enters the stomach, HCl is released and the \([H_3O^+]\) of the stomach fluid rises to \(4 \times 10^{-2}\) M. What is the pH of the stomach fluid while eating?

14.83 In Larry’s esophageal pH test, a pH value of 3.60 was recorded in the esophagus. What is the \([H_3O^+]\) in his esophagus?

14.84 After Larry had taken Nexium for 4 weeks, the pH in his stomach was raised to 4.52. What is the \([H_3O^+]\) in his stomach?

14.85 Write the balanced chemical equation for the neutralization reaction of stomach acid HCl with CaCO₃, an ingredient in some antacids.

14.86 Write the balanced chemical equation for the neutralization reaction of stomach acid HCl with Al(OH)₃, an ingredient in some antacids.

14.87 How many grams of CaCO₃ are required to neutralize 100. mL of stomach acid HCl, which is equivalent to 0.0400 M HCl?

14.88 How many grams of Al(OH)₃ are required to neutralize 150. mL of stomach acid HCl with a pH of 1.5?
14.1 Acids and Bases

**LEARNING GOAL** Describe and name acids and bases.

- An Arrhenius acid produces H\(^+\) and an Arrhenius base produces OH\(^-\) in aqueous solutions.
- Acids taste sour, may sting, and neutralize bases.
- Bases taste bitter, feel slippery, and neutralize acids.
- Acids containing a simple anion use a hydro prefix, whereas acids with oxygen-containing polyatomic anions are named as ic or ous acids.

14.2 Brønsted–Lowry Acids and Bases

**LEARNING GOAL** Identify conjugate acid–base pairs for Brønsted–Lowry acids and bases.

According to the Brønsted–Lowry theory, acids are H\(^+\) donors and bases are H\(^+\) acceptors.
- A conjugate acid–base pair is related by the loss or gain of one H\(^+\).
- For example, when the acid HF donates H\(^+\), the F\(^-\) is its conjugate base. The other acid–base pair would be H\(_3\text{O}^+\)/H\(_2\text{O}\).

14.3 Strengths of Acids and Bases

**LEARNING GOAL** Write equations for the dissociation of strong and weak acids; identify the direction of reaction.

- Strong acids dissociate completely in water, and the H\(^+\) is accepted by H\(_2\text{O}\) acting as a base.
- A weak acid dissociates slightly in water, producing only a small percentage of H\(_3\text{O}^+\).
- Strong bases are hydroxides of Groups 1A (1) and 2A (2) that dissociate completely in water.
- An important weak base is ammonia, NH\(_3\).
14.4 Dissociation Constants for Acids and Bases

LEARNING GOAL
Write the dissociation expression for a weak acid or weak base.

- In water, weak acids and weak bases produce only a few ions when equilibrium is reached.
- Weak acids have small $K_a$ values whereas strong acids, which are essentially 100% dissociated, have very large $K_a$ values.
- The reaction for a weak acid can be written as $HA + H_2O \rightleftharpoons H_3O^+ + A^-$. The acid dissociation expression is written as
  $$K_a = \frac{[H_3O^+][A^-]}{[HA]}.$$ 
- For a weak base, $B + H_2O \rightleftharpoons BH^+ + OH^-$, the base dissociation expression is written as
  $$K_b = \frac{[BH^+][OH^-]}{[B]}.$$ 

14.5 Dissociation of Water

LEARNING GOAL
Use the water dissociation expression to calculate the $[H_2O^+]$ and $[OH^-]$ in an aqueous solution.

- In pure water, a few water molecules transfer $H^+$ to other water molecules, producing small, but equal, amounts of $[H_2O^+]$ and $[OH^-]$.
- In pure water, the molar concentrations of $H_2O^+$ and $OH^-$ are each $1.0 \times 10^{-7}$ mol/L.
- The water dissociation expression, $K_w$, $[H_2O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 $^\circ$C.
- In acidic solutions, the $[H_2O^+]$ is greater than the $[OH^-]$.
- In basic solutions, the $[OH^-]$ is greater than the $[H_2O^+]$.

14.6 The pH Scale

LEARNING GOAL
Calculate pH from $[H_2O^+]$; given the pH, calculate the $[H_2O^+]$ and $[OH^-]$ of a solution.

- The pH scale is a range of numbers typically from 0 to 14, which represents the $[H_2O^+]$ of the solution.
- A neutral solution has a pH of 7.0. In acidic solutions, the pH is below 7.0; in basic solutions, the pH is above 7.0.
- Mathematically, pH is the negative logarithm of the hydronium ion concentration, $pH = -\log[H_2O^+]$.

14.7 Reactions of Acids and Bases

LEARNING GOAL
Write balanced equations for reactions of acids with metals, carbonates, or bicarbonates, and bases.

- An acid reacts with a metal to produce hydrogen gas and a salt.
- The reaction of an acid with a carbonate or bicarbonate produces carbon dioxide, water, and a salt.
- In neutralization, an acid reacts with a base to produce water and a salt.

14.8 Acid–Base Titration

LEARNING GOAL
Calculate the molarity or volume of an acid or base solution from titration information.

- In a titration, an acid sample is neutralized with a known amount of a base.
- From the volume and molarity of the base, the concentration of the acid is calculated.

14.9 Buffers

LEARNING GOAL
Describe the role of buffers in maintaining the pH of a solution; calculate the pH of a buffer.

- A buffer solution resists changes in pH when small amounts of an acid or a base are added.
- A buffer contains either a weak acid and its salt or a weak base and its salt.
- In a buffer, the weak acid reacts with added $OH^-$, and the anion of the salt reacts with added $H_2O^+$.
- Most buffer solutions consist of nearly equal concentrations of a weak acid and a salt containing its conjugate base.
- The pH of a buffer is calculated by solving the $K_a$ expression for $[H_2O^+]$.

KEY TERMS

- **acid**: A substance that dissolves in water and produces hydrogen ions ($H^+$), according to the Arrhenius theory. All acids are hydrogen ion donors, according to the Brønsted–Lowry theory.
- **acid dissociation expression, $K_a$**: The product of the ions from the dissociation of a weak acid divided by the concentration of the weak acid.
- **amphoteric**: Substances that can act as either an acid or a base in water.
- **base**: A substance that dissolves in water and produces hydroxide ions ($OH^-$) according to the Arrhenius theory. All bases are hydrogen ion acceptors, according to the Brønsted–Lowry theory.
- **base dissociation expression, $K_b$**: The product of the ions from the dissociation of a weak base divided by the concentration of the weak base.
**Brønsted–Lowry acids and bases** An acid is a hydrogen ion donor; a base is a hydrogen ion acceptor.

**buffer solution** A solution of a weak acid and its conjugate base or a weak base and its conjugate acid that maintains the pH by neutralizing added acid or base.

**conjugate acid–base pair** An acid and a base that differ by one H⁺.

When an acid donates a hydrogen ion, the product is its conjugate base, which is capable of accepting a hydrogen ion in the reverse reaction.

**dissociation** The separation of an acid or a base into ions in water.

**endpoint** The point at which an indicator changes color. For the indicator phenolphthalein, the color change occurs when the number of moles of OH⁻ is equal to the number of moles of H₃O⁺ in the sample.

**hydronium ion, H₃O⁺** The ion formed by the attraction of a hydrogen ion, H⁺, to a water molecule.

**indicator** A substance added to a titration sample that changes color when the pH of the solution changes.

---

**KEY MATH SKILLS**

The chapter section containing each Key Math Skill is shown in parentheses at the end of each heading.

**Calculating pH from [H₃O⁺] (14.6)**

- The pH of a solution is calculated from the negative log of the [H₃O⁺].

  \[ \text{pH} = -\log([\text{H}_3\text{O}^+] \] \[= -\log(2.4 \times 10^{-11} \text{ M}) \] 

**Example:** What is the pH of a solution that has [H₃O⁺] = 2.4 × 10⁻¹¹ M?

**Answer:** We substitute the given [H₃O⁺] into the pH equation and calculate the pH.

  \[ \text{pH} = -\log([\text{H}_3\text{O}^+] \] \[= -\log(2.4 \times 10^{-11} \text{ M}) \] 

  = 10.62 Two decimal places equal the two SFs in the [H₃O⁺] coefficient.

**CORE CHEMISTRY SKILLS**

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

**Identifying Conjugate Acid–Base Pairs (14.2)**

- According to the Brønsted–Lowry theory, a conjugate acid–base pair consists of molecules or ions related by the loss of one H⁺ by an acid, and the gain of one H⁺ by a base.

- Every acid–base reaction contains two conjugate acid–base pairs because an H⁺ is transferred in both the forward and reverse directions.

- When an acid such as HF loses one H⁺, the conjugate base F⁻ is formed. When H₂O acts as a base, it gains one H⁺, which forms its conjugate acid, H₃O⁺.

**Example:** Identify the conjugate acid–base pairs in the following reaction:

\[ \text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow \text{HSO}_4^-(aq) + \text{H}_3\text{O}^+(aq) \]
• If we know the $[\text{H}_2\text{O}^+]$ of a solution, we can use the $K_w$ expression to calculate the $[\text{OH}^-]$. If we know the $[\text{OH}^-]$ of a solution, we can calculate the $[\text{H}_2\text{O}^+]$ using the $K_w$ expression.

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_2\text{O}^+]} \quad \text{and} \quad [\text{H}_2\text{O}^+] = \frac{K_w}{[\text{OH}^-]}$$

**Example:** What is the $[\text{OH}^-]$ in a solution that has $[\text{H}_2\text{O}^+] = 2.4 \times 10^{-14}$ M? Is the solution acidic or basic?

**Answer:** We solve the $K_w$ expression for $[\text{OH}^-]$ and substitute in the known values of $K_w$ and $[\text{H}_2\text{O}^+]$.

$$[\text{OH}^-] = \frac{2.4 \times 10^{-14}}{[\text{H}_2\text{O}^+]} = 4.2 \times 10^{-4} \text{ M}$$

Because the $[\text{OH}^-]$ is greater than the $[\text{H}_2\text{O}^+]$, this is a basic solution.

**Writing Equations for Reactions of Acids and Bases (14.7)**

• Acids react with certain metals to produce hydrogen gas ($\text{H}_2$) and a salt.

$$\text{Mg}(s) + 2\text{HCl}(aq) \rightarrow \text{H}_2(g) + \text{MgCl}_2(aq)$$

**Metal** | **Acid** | **Hydrogen** | **Salt**
---|---|---|---

• When an acid is added to a carbonate or bicarbonate, the products are carbon dioxide gas, water, and a salt.

$$2\text{HCl}(aq) + \text{Na}_2\text{CO}_3(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + 2\text{NaCl}(aq)$$

**Acid** | **Carbonate** | **Carbon** | **Water** | **Salt**
---|---|---|---|---

• Neutralization is a reaction between a strong or weak acid and a strong base to produce water and a salt.

$$\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq)$$

**Acid** | **Base** | **Water** | **Salt**
---|---|---|---

**Example:** Write the balanced chemical equation for the reaction of ZnCO$_3$(s) and hydrobromic acid HBr(aq).

**Answer:**

$$\text{ZnCO}_3(s) + 2\text{HBr}(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{ZnBr}_2(aq)$$

**Calculating Molarity or Volume of an Acid or Base in a Titration (14.8)**

• In a titration, a measured volume of acid is neutralized by a NaOH solution of known molarity.

• From the measured volume of the NaOH solution required for titration and its molarity, the number of moles of NaOH, the moles of acid, and the concentration of the acid are calculated.

**Example:** A 15.0-mL sample of a H$_2$SO$_4$ solution is titrated with 24.0 mL of a 0.245 M NaOH solution. What is the molarity of the H$_2$SO$_4$ solution?

$$\text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{Na}_2\text{SO}_4(aq)$$

**Answer:**

$$\text{Molarity (M)} = \frac{0.00294 \text{ mol H}_2\text{SO}_4}{0.0150 \text{ L NaOH solution}} = 0.196 \text{ M H}_2\text{SO}_4$$

**Calculating the pH of a Buffer (14.9)**

• A buffer solution maintains pH by neutralizing small amounts of added acid or base.

• Most buffer solutions consist of nearly equal concentrations of a weak acid and a salt containing its conjugate base such as acetic acid, HC$_2$H$_3$O$_2$ and its salt Na$_2$C$_2$H$_3$O$_2$.

• The $[\text{H}_3\text{O}^+]$ is calculated by solving the $K_a$ expression for $[\text{H}_3\text{O}^+]$, then substituting the values of $[\text{H}_3\text{O}^+]$, $[\text{HA}]$, and $K_a'$ into the equation.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

Solving for $[\text{H}_3\text{O}^+]$ gives:

$$[\text{H}_3\text{O}^+] = K_a \times \frac{[\text{HC}_2\text{H}_3\text{O}_2]}{[\text{C}_2\text{H}_3\text{O}_2^-]}$$

**Example:** What is the pH of a buffer prepared with 0.40 M HC$_2$H$_3$O$_2$ and 0.20 M C$_2$H$_3$O$_2^-$ if the $K_a$ of acetic acid is $1.8 \times 10^{-5}$?

**Answer:**

$$[\text{H}_3\text{O}^+] = K_a \times \frac{[\text{HC}_2\text{H}_3\text{O}_2]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = 1.8 \times 10^{-5} \times \frac{0.40}{0.20} = 3.6 \times 10^{-5} \text{ M}$$

$$\text{pH} = -\log[3.6 \times 10^{-5}] = 4.44$$

**UNDERSTANDING THE CONCEPTS**

*The chapter sections to review are shown in parentheses at the end of each question.*

**14.89** Determine if each of the following diagrams represents a strong acid or a weak acid. The acid has the formula HX. (14.3)

**14.90** Adding a few drops of a strong acid to water will lower the pH appreciably. However, adding the same number of drops to a buffer does not appreciably alter the pH. Why? (14.9)
14.94 Complete the following table: (14.2)

<table>
<thead>
<tr>
<th>Base</th>
<th>Conjugate Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>HS⁻</td>
<td>H₂O⁺</td>
</tr>
<tr>
<td>NH₃</td>
<td>HCO₃⁻</td>
</tr>
</tbody>
</table>

Applications

14.95 Sometimes, during stress or trauma, a person can start to hyperventilate. Then the person might breathe into a paper bag to avoid fainting. (14.9)
a. What changes occur in the blood pH during hyperventilation?

**Additional Questions and Problems**

14.99 Identify each of the following as an acid, base, or salt, and give its name: (14.1)
a. HBrO₂  b. CsOH  c. Mg(NO₃)₂  d. HClO₄
14.100 Identify each of the following as an acid, base, or salt, and give its name: (14.1)
a. HNO₂  b. MgBr₂  c. NH₃  d. Li₂SO₄
14.101 Complete the following table: (14.2)

<table>
<thead>
<tr>
<th>Acid</th>
<th>Conjugate Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>HI</td>
<td>Cl⁻</td>
</tr>
<tr>
<td>NH₃⁺</td>
<td>HS⁻</td>
</tr>
</tbody>
</table>

14.102 Complete the following table: (14.2)

<table>
<thead>
<tr>
<th>Base</th>
<th>Conjugate Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>PH₃</td>
<td>HClO₄</td>
</tr>
<tr>
<td></td>
<td>CH₃OH</td>
</tr>
</tbody>
</table>

14.103 Using Table 14.3, identify the stronger acid in each of the following pairs: (14.3)
a. HI or NH₄⁺  b. HClO₄ or H₂S  c. HNO₂ or HCl₂H₃O₂  d. H₂O or HCO₃⁻
14.104 Using Table 14.3, identify the stronger acid in each of the following pairs: (14.3)
a. HI or NH₄⁺  b. HClO₄ or H₂S  c. HNO₂ or HCl₂H₃O₂  d. H₂O or HCO₃⁻
14.105 Determine the pH and pOH for each of the following solutions: (14.6)
a. [H₂O⁺] = 2.0 × 10⁻⁸ M  b. [H₂O⁺] = 5.0 × 10⁻² M  c. [OH⁻] = 3.5 × 10⁻⁴ M  d. [OH⁻] = 0.0054 M
14.106 Determine the pH and pOH for each of the following solutions: (14.6)
a. [OH⁻] = 1.0 × 10⁻⁷ M  b. [H₂O⁺] = 4.2 × 10⁻³ M  c. [H₂O⁺] = 0.0001 M  d. [OH⁻] = 8.5 × 10⁻⁹ M
14.107 Are the solutions in problem 14.105 acidic, basic, or neutral? (14.6)
14.108 Are the solutions in problem 14.106 acidic, basic, or neutral? (14.6)
14.109 Calculate the [H₂O⁺] and [OH⁻] for a solution with each of the following pH values: (14.6)
a. 4.5  b. 6.9  c. 8.1  d. 5.8
14.110 Calculate the [H₂O⁺] and [OH⁻] for a solution with each of the following pH values: (14.6)
a. 2.2  b. 7.2  c. 10.6  d. 12.1
14.114 What is the pH and pOH of a solution prepared by dissolving 1.0 g of Ca(OH)$_2$ in water to make 875 mL of solution? (14.6)

14.123 One of the most acidic lakes in the United States is Little Echo Pond in the Adirondacks in New York. Recently, this lake had a pH of 4.2, well below the recommended pH of 6.5. (14.6, 14.8)
a. What are the [H$_3$O$^+$] and [OH$^-$] of Little Echo Pond?
b. What are the [H$_3$O$^+$] and [OH$^-$] of a lake that has a pH of 6.5?
c. One way to raise the pH of an acidic lake (and restore aquatic life) is to add limestone (CaCO$_3$). How many grams of CaCO$_3$ are needed to neutralize 1.0 kL of the acidic water from the lake if the acid is sulfuric acid? H$_2$SO$_4$(aq) + CaCO$_3$(s) $\longrightarrow$ CO$_2$(g) + H$_2$O(l) + CaSO$_4$(aq)

A helicopter drops calcium carbonate on an acidic lake to increase its pH.

Applications

14.124 The daily output of stomach acid (gastric juice) is 1000 mL to 2000 mL. Prior to a meal, stomach acid (HCl) typically has a pH of 1.42. (14.6, 14.7, 14.8)
a. What is the [H$_3$O$^+$] of stomach acid?
b. One chewable tablet of the antacid Maalox contains 600. mg of CaCO$_3$. Write the neutralization equation, and calculate the milliliters of stomach acid neutralized by two tablets of Maalox.
c. The antacid milk of magnesia contains 400. mg of Mg(OH)$_2$ per teaspoon. Write the neutralization equation, and calculate the number of milliliters of stomach acid that are neutralized by 1 tablespoon of milk of magnesia.

14.125 Calculate the volume, in milliliters, of a 0.150 M NaOH solution that will completely neutralize each of the following: (14.8)
a. 25.0 mL of a 0.288 M HCl solution
b. 10.0 mL of a 0.560 M H$_2$SO$_4$ solution
14.126 Calculate the volume, in milliliters, of a 0.215 M NaOH solution that will completely neutralize each of the following: (14.8)
   a. 3.80 mL of a 1.25 M HNO₃ solution
   b. 8.50 mL of a 0.825 M H₃PO₄ solution

14.127 A solution of 0.205 M NaOH is used to reach 20.0 mL of a H₂SO₄ solution. If 45.6 mL of the NaOH solution is required to reach the endpoint, what is the molarity of the H₂SO₄ solution? (14.8)
   \[ \text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{Na}_2\text{SO}_4(aq) \]

14.128 A 10.0-mL sample of vinegar, which is an aqueous solution of acetic acid, H₃C₂H₃O₂, requires 16.5 mL of a 0.20 M HCl solution to reach the endpoint in a titration. What is the molarity of the acetic acid solution? (14.8)
   \[ \text{HC}_2\text{H}_3\text{O}_2(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaC}_2\text{H}_3\text{O}_2(aq) \]

14.129 A buffer solution is made by dissolving H₃PO₄ and NaH₂PO₄ in water. (14.9)
   a. Write an equation that shows how this buffer neutralizes added acid.
   b. Write an equation that shows how this buffer neutralizes added base.
   c. Calculate the pH of this buffer if it contains 0.50 M H₃PO₄ and 0.20 M H₂PO₄⁻. The \( K_a \) for H₃PO₄ is \( 7.5 \times 10^{-2} \).

14.130 A buffer solution is made by dissolving HC₂H₃O₂ and NaC₂H₃O₂ in water. (14.9)
   a. Write an equation that shows how this buffer neutralizes added acid.
   b. Write an equation that shows how this buffer neutralizes added base.
   c. Calculate the pH of this buffer if it contains 0.20 M HC₂H₃O₂ and 0.40 M C₂H₃O₂⁻. The \( K_a \) for HC₂H₃O₂ is \( 1.8 \times 10^{-5} \).

14.23 a. reactants   b. reactants   c. products
14.25 \( \text{NH}_4^+(aq) + \text{SO}_4^{2-}(aq) \rightleftharpoons \text{NH}_3(aq) + \text{HSO}_4^-(aq) \)

   The equilibrium mixture contains mostly reactants because NH₄⁺ is a weaker acid than HSO₄⁻, and SO₄²⁻ is a weaker base than NH₃.

14.27 a. true   b. false   c. false
   d. true   e. false

14.29 a. H₂SO₃   b. HSO₃⁻   c. H₂SO₄
   d. HS⁻   d. H₂SO₃

14.31 \( \text{H}_3\text{PO}_4(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{H}_2\text{PO}_4^-(aq) \)

   \[ K_a = \frac{[\text{H}_3\text{O}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]} \]

14.33 In pure water, [H₃O⁺] = [OH⁻] because one of each is produced every time a hydrogen ion is transferred from one water molecule to another.

14.35 In an acidic solution, the [H₃O⁺] is greater than the [OH⁻].

14.37 a. acidic   b. basic
   c. basic   d. acidic

14.39 a. 1.0 \times 10^{-5} M   b. 1.0 \times 10^{-8} M
   c. 5.0 \times 10^{-10} M   d. 2.5 \times 10^{-2} M

14.41 a. 2.5 \times 10^{-13} M   b. 2.0 \times 10^{-9} M
   c. 5.0 \times 10^{-11} M   d. 1.3 \times 10^{-6} M

14.43 In a neutral solution, the [H₃O⁺] is 1.0 \times 10^{-7} M and the pH is 7.00, which is the negative value of the power of 10.

14.45 a. basic   b. acidic   c. basic
   d. acidic   e. acidic   f. basic

14.47 An increase or decrease of one pH unit changes the [H₃O⁺] by a factor of 10. Thus a pH of 3 is 10 times more acidic than a pH of 4.

14.49 a. 4.0   b. 8.5   c. 9.0
   d. 3.40   e. 7.17   f. 10.92
14.51

<table>
<thead>
<tr>
<th>[H$_3$O$^+$]</th>
<th>[OH$^-$]</th>
<th>pH</th>
<th>pOH</th>
<th>Acidic, Basic, or Neutral?</th>
</tr>
</thead>
<tbody>
<tr>
<td>$1.0 \times 10^{-8}$ M</td>
<td>$1.0 \times 10^{-6}$ M</td>
<td>8.00</td>
<td>6.00</td>
<td>Basic</td>
</tr>
<tr>
<td>$3.2 \times 10^{-4}$ M</td>
<td>$3.1 \times 10^{-11}$ M</td>
<td>3.49</td>
<td>10.51</td>
<td>Acidic</td>
</tr>
<tr>
<td>$2.8 \times 10^{-3}$ M</td>
<td>$3.6 \times 10^{-10}$ M</td>
<td>4.55</td>
<td>9.45</td>
<td>Acidic</td>
</tr>
<tr>
<td>$1.0 \times 10^{-12}$ M</td>
<td>$1.0 \times 10^{-2}$ M</td>
<td>12.00</td>
<td>2.00</td>
<td>Basic</td>
</tr>
</tbody>
</table>

14.53 $1.2 \times 10^{-7}$ M

14.55 a. ZnCO$_3$(s) + 2HBr(aq) $\rightarrow$ CO$_2$(g) + H$_2$O(l) + ZnBr$_2$(aq)
b. Zn(s) + 2HCl(aq) $\rightarrow$ H$_2$(g) + ZnCl$_2$(aq)
c. HCl(aq) + NaHCO$_3$(s) $\rightarrow$ CO$_2$(g) + H$_2$O(l) + NaCl(aq)
d. H$_2$SO$_4$(aq) + Mg(OH)$_2$(s) $\rightarrow$ 2H$_2$O(l) + MgSO$_4$(aq)

14.57 a. 2HCl(aq) + Mg(OH)$_2$(s) $\rightarrow$ 2H$_2$O(l) + MgCl$_2$(aq)
b. H$_2$PO$_4$(aq) + 3LiOH(aq) $\rightarrow$ 3H$_2$O(l) + Li$_3$PO$_4$(aq)

14.59 a. H$_2$SO$_4$(aq) + 2NaOH(aq) $\rightarrow$ 2H$_2$O(l) + Na$_2$SO$_4$(aq)
b. 3HCl(aq) + Fe(OH)$_3$(s) $\rightarrow$ 3H$_2$O(l) + FeCl$_3$(aq)
c. H$_2$CO$_3$(aq) + Mg(OH)$_2$(s) $\rightarrow$ 2H$_2$O(l) + MgCO$_3$(s)

14.61 To a known volume of a formic acid solution, add a few drops of indicator. Place a solution of NaOH of known molarity in a buret. Add base to acid until one drop changes the color of the solution. Use the volume and molarity of NaOH and the volume of a formic acid solution to calculate the concentration of the formic acid in the sample.

14.63 0.829 M HCl solution
14.65 0.124 M H$_2$SO$_4$ solution
14.67 16.5 mL

14.69 b and c are buffer systems. b contains the weak acid H$_2$CO$_3$ and its salt NaHCO$_3$. c contains HF, a weak acid, and its salt KF.

14.71 a. 3  b. 1 and 2  c. 3  d. 1

14.73 pH = 3.35

14.75 The pH of the 0.10 M HF/0.10 M NaF buffer is 3.46.
The pH of the 0.060 M HF/0.120 M NaF buffer is 3.76.

14.77 If you breathe fast, CO$_2$ is expelled and the equilibrium shifts to lower H$_2$O$^+$, which raises the pH.

14.79 If large amounts of HCO$_3^-$ are lost, equilibrium shifts to higher H$_2$O$^+$, which lowers the pH.

14.81 pH = 3.70

14.83 $2.5 \times 10^{-4}$ M
14.85 CaCO$_3$(s) + 2HCl(aq) $\rightarrow$ CO$_2$(g) + H$_2$O(l) + CaCl$_2$(aq)
14.87 0.200 g of CaCO$_3$

14.89 a. This diagram represents a weak acid; only a few HX molecules separate into H$_2$O$^+$ and X$^-$ ions.
b. This diagram represents a strong acid; all the HX molecules separate into H$_2$O$^+$ and X$^-$ ions.

14.91 a. base  b. acid  c. acid  d. base

14.93

<table>
<thead>
<tr>
<th>Acid</th>
<th>Conjugate Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>H$_2$SO$_3$</td>
<td>HSO$_3^-$</td>
</tr>
<tr>
<td>CH$_3$CH$_2$COOH</td>
<td>CH$_3$CH$_2$COO$^-$</td>
</tr>
<tr>
<td>NH(CH$_3$)$_3^+$</td>
<td>N(CH$_3$)$_3$</td>
</tr>
<tr>
<td>CC$_2$COOH</td>
<td>CC$_2$COO$^-$</td>
</tr>
</tbody>
</table>

14.95 a. During hyperventilation, a person will lose CO$_2$ and the blood pH will rise.
b. Breathing into a paper bag will increase the CO$_2$ concentration and lower the blood pH.

14.97 a. acidic  b. acidic  c. neutral  d. basic

14.99 a. acid, bromous acid  b. base, cesium hydroxide  c. salt, magnesium nitrate  d. acid, perchloric acid

14.101

<table>
<thead>
<tr>
<th>Acid</th>
<th>Conjugate Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>HI</td>
<td>I$^-$</td>
</tr>
<tr>
<td>HCl</td>
<td>Cl$^-$</td>
</tr>
<tr>
<td>NH$_4^+$</td>
<td>NH$_3$</td>
</tr>
<tr>
<td>H$_2$S</td>
<td>HS$^-$</td>
</tr>
</tbody>
</table>

14.103 a. HF  b. H$_2$O$^+$  c. HNO$_2$  d. HCO$_3^-$

14.105 a. pH = 7.70; pOH = 6.30  b. pH = 1.30; pOH = 12.70  c. pH = 10.54; pOH = 3.46  d. pH = 11.73; pOH = 2.27

14.107 a. basic  b. acidic  c. basic  d. basic


14.113 pH = 1.04; pOH = 12.9

14.115 a. 1. NO$_3^-$
2. F$^-$
b. 1. [H$_2$O$^+$][NO$_3^-$] / [HNO$_3$]  
2. [H$_2$O$^+$][F$^-$] / [HF]
c. HF is the weaker acid.

14.117 a. HNO$_3$/NO$_3^-$ and NH$_4^+$/NH$_3$; equilibrium mixture contains mostly products  
b. HBr/Br$^-$ and H$_2$O$^+$/H$_2$O; equilibrium mixture contains mostly products
14.119  a. \( \text{HCl}(aq) + \text{LiOH}(s) \rightarrow \text{LiCl}(aq) + \text{H}_2\text{O}(l) \)
    b. \( \text{MgCO}_3(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{MgSO}_4(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l) \)

14.121  a. \([\text{H}_3\text{O}^+]=2.0 \times 10^{-13}\text{ M}\)
    b. pH = 12.70
    c. pOH = 1.30
    d. \(2\text{KOH}(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{K}_2\text{SO}_4(aq)\)
    e. 56 mL of the KOH solution

14.123  a. \([\text{H}_3\text{O}^+] = 6 \times 10^{-5}\text{ M}; [\text{OH}^-] = 2 \times 10^{-10}\text{ M}\)
    b. \([\text{H}_3\text{O}^+] = 3 \times 10^{-7}\text{ M}; [\text{OH}^-] = 3 \times 10^{-8}\text{ M}\)
    c. 3 g of CaCO\(_3\)

14.125  a. 48.0 mL of NaOH solution
    b. 74.7 mL of NaOH solution

14.127  0.234 M \(\text{H}_2\text{SO}_4\) solution

14.129  a. acid:
    \(\text{H}_3\text{PO}_4(aq) + \text{H}_3\text{O}^+(aq) \rightarrow \text{H}_3\text{PO}_4(aq) + \text{H}_2\text{O}(l)\)
    b. base:
    \(\text{H}_3\text{PO}_4(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{PO}_4^-(aq) + \text{H}_2\text{O}(l)\)
    c. pH = 1.72
CI.21 Methane is a major component of purified natural gas used for heating and cooking. When 1.0 mol of methane gas burns with oxygen to produce carbon dioxide and water vapor, 883 kJ of heat is produced. At STP, methane gas has a density of 0.715 g/L. For transport, the natural gas is cooled to −163 °C to form liquefied natural gas (LNG) with a density of 0.45 g/mL. A tank on a ship can hold 7.0 million gallons of LNG. (2.7, 7.2, 7.3, 8.2, 8.3, 9.5, 10.1, 11.6)

An LNG carrier transports liquefied natural gas.

a. Draw the Lewis structure for methane, which has the formula CH₄.
b. What is the mass, in kilograms, of LNG (assume that LNG is all methane) transported in one tank on a ship?
c. What is the volume, in liters, of LNG (methane) from one tank when the LNG (methane) from one tank is converted to methane gas at STP?
d. Write the balanced chemical equation for the combustion of methane and oxygen in a gas burner, including the heat of reaction.

two gases found in automobile exhaust are carbon dioxide and nitrogen oxide.

Methane is the fuel burned in a gas cooktop.

e. How many kilograms of oxygen are needed to react with all of the methane in one tank of LNG?
f. How much heat, in kilojoules, is released after burning all of the methane from one tank of LNG?

CI.22 Automobile exhaust is a major cause of air pollution. One pollutant is nitrogen oxide, which forms from nitrogen and oxygen gases in the air at the high temperatures in an automobile engine. Once emitted into the air, nitrogen oxide reacts with oxygen to produce nitrogen dioxide, a reddish brown gas with a sharp, pungent odor that makes up smog. One component of gasoline is octane, C₈H₁₈, which has a density of 0.803 g/mL. In one year, a typical automobile uses 550 gal of gasoline and produces 41 lb of nitrogen oxide. (2.7, 7.2, 7.3, 8.2, 9.2, 11.6)

a. Write balanced chemical equations for the production of nitrogen oxide and nitrogen dioxide.
b. If all the nitrogen oxide emitted by one automobile is converted to nitrogen dioxide in the atmosphere, how many kilograms of nitrogen dioxide are produced in one year by a single automobile?
c. Write a balanced chemical equation for the combustion of octane.
d. How many moles of C₈H₁₈ are present in 15.2 gal of octane?
e. How many liters of CO₂ at STP are produced in one year from the gasoline used by the typical automobile?

CI.23 A mixture of 25.0 g of CS₂ gas and 30.0 g of O₂ gas is placed in a 10.0-L container and heated to 125 °C. The products of the reaction are carbon dioxide gas and sulfur dioxide gas. (7.2, 7.3, 8.2, 9.3, 11.1, 11.7)

a. Write a balanced chemical equation for the reaction.
b. How many grams of CO₂ are produced?
c. What is the partial pressure, in millimeters of mercury, of the remaining reactant?
d. What is the final pressure, in millimeters of mercury, in the container?

CI.24 In wine-making, glucose (C₆H₁₂O₆) from grapes undergoes fermentation in the absence of oxygen to produce ethanol and carbon dioxide. A bottle of vintage port wine has a volume of 750 mL and contains 135 mL of ethanol (C₂H₆O). Ethanol has a density of 0.789 g/mL. In 1.5 lb of grapes, there are 26 g of glucose. (2.7, 7.2, 7.3, 8.2, 9.2, 12.4)

a. Calculate the volume percent (v/v) of ethanol in the port wine.
b. What is the molarity (M) of ethanol in the port wine?
c. Write the balanced chemical equation for the fermentation reaction of sugar in grapes.
d. How many grams of sugar from grapes are required to produce one bottle of port wine?
e. How many bottles of port wine can be produced from 1.0 ton of grapes (1 ton = 2000 lb)?

Port is a type of fortified wine that is produced in Portugal.

When the glucose in grapes is fermented, ethanol is produced.
CI.25 Consider the following reaction at equilibrium:

\[ 2\text{H}_2(g) + \text{S}_2(g) \rightleftharpoons 2\text{H}_2\text{S}(g) + \text{heat} \]

In a 10.0-L container, an equilibrium mixture contains 2.02 g of \( \text{H}_2 \), 10.3 g of \( \text{S}_2 \), and 68.2 g of \( \text{H}_2\text{S} \). (7.2, 7.3, 13.2, 13.3, 13.4, 13.5)

a. What is the numerical value of \( K_c \) for this equilibrium mixture?
b. If more \( \text{H}_2 \) is added to the equilibrium mixture, how will the equilibrium shift?
c. How will the equilibrium shift if the mixture is placed in a 5.00-L container with no change in temperature?
d. If a 5.00-L container has an equilibrium mixture of 0.300 mol of \( \text{H}_2 \) and 2.50 mol of \( \text{H}_2\text{S} \), what is the [\( \text{S}_2 \)] if temperature remains constant?

CI.26 A saturated solution of silver hydroxide has a pH of 10.15. (7.2, 7.3, 13.2, 13.6)

a. Write the solubility product expression for silver hydroxide.
b. Calculate the numerical value of \( K_{sp} \) for silver hydroxide.
c. How many grams of silver hydroxide will dissolve in 2.0 L of water?

d. How many milliliters of the HCl solution are needed to neutralize the \( \text{Mg(OH)}_2 \)?
e. How many milliliters of the HCl solution are needed to neutralize the \( \text{Al(OH)}_3 \)?

CI.29 A KOH solution is prepared by dissolving 8.57 g of KOH in enough water to make 850. mL of KOH solution. (12.4, 14.6, 14.7, 14.8)

a. What is the molarity of the KOH solution?
b. What is the [\( \text{H}_3\text{O}^+ \)] and pH of the KOH solution?
c. Write the balanced chemical equation for the neutralization of KOH by \( \text{H}_2\text{SO}_4 \).
d. How many milliliters of a 0.250 M \( \text{H}_2\text{SO}_4 \) solution is required to neutralize 10.0 mL of the KOH solution?

CI.30 A solution of HCl is prepared by diluting 15.0 mL of a 12.0 M HCl solution with enough water to make 750. mL of HCl solution. (12.4, 14.6, 14.7, 14.8)

a. What is the molarity of the HCl solution?
b. What is the [\( \text{H}_3\text{O}^+ \)] and pH of the HCl solution?
c. Write the balanced chemical equation for the reaction of HCl and \( \text{MgCO}_3 \).
d. How many milliliters of the diluted HCl solution is required to completely react with 350. mg of \( \text{MgCO}_3 \)?

c. Write a balanced chemical equation that shows how this buffer neutralizes added base.
f. Write a balanced chemical equation that shows how this buffer neutralizes added acid.

CI.31 A volume of 200.0 mL of a carbonic acid buffer for blood plasma is prepared that contains 0.403 g of \( \text{NaHCO}_3 \) and 0.0149 g of \( \text{H}_2\text{CO}_3 \). At body temperature (37 °C), the \( K_a \) of carbonic acid is \( 7.9 \times 10^{-7} \). (7.2, 7.3, 8.4, 14.9)

\[ \text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCO}_3^{-}(aq) + \text{H}_3\text{O}^+(aq) \]

a. What is the [\( \text{H}_2\text{CO}_3 \)]?
b. What is the [\( \text{HCO}_3^{-} \)]?
c. What is the [\( \text{H}_3\text{O}^+ \)]?
d. What is the pH of the buffer?
e. Write a balanced chemical equation that shows how this buffer neutralizes added acid.
f. Write a balanced chemical equation that shows how this buffer neutralizes added base.

c. Write a balanced chemical equation that shows how this buffer neutralizes added base.

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CI.32 In the kidneys, the ammonia buffer system buffers high \( \text{H}_2\text{O}^+ \). Ammonia, which is produced in renal tubules from amino acids, combines with \( \text{H}^+ \) to be excreted as \( \text{NH}_4\text{Cl} \). At body temperature (37 °C), the \( K_a \) of ammonia is \( 5.6 \times 10^{-10} \). A buffer solution with a volume of 125 mL contains 3.34 g of \( \text{NH}_4\text{Cl} \) and 0.0151 g of \( \text{NH}_3 \). (7.2, 7.3, 8.4, 14.9)

\[ \text{NH}_4^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_3(aq) + \text{H}_3\text{O}^+(aq) \]

a. What is the [\( \text{NH}_4^+ \)]?
b. What is the [\( \text{NH}_3 \)]?
c. What is the [\( \text{H}_3\text{O}^+ \)]?
d. What is the pH of the buffer?
e. Write a balanced chemical equation that shows how this buffer neutralizes added acid.
f. Write a balanced chemical equation that shows how this buffer neutralizes added base.

c. Write a balanced chemical equation that shows how this buffer neutralizes added base.

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c. Write a balanced chemical equation that shows how this buffer neutralizes added base.
ANSWERS

CI.21  
a. $\text{H} \cdot \text{C} \cdot \text{H}$ or $\text{H} - \text{C} - \text{H}$  

b. $1.2 \times 10^7$ kg of LNG (methane)  
c. $1.7 \times 10^{10}$ L of LNG (methane)  
d. $\text{CH}_4(g) + 2\text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + 2\text{H}_2\text{O}(g) + 883 \text{ kJ}$  
e. $4.8 \times 10^7$ kg of $\text{O}_2$  
f. $6.6 \times 10^{11} \text{ kJ}$

CI.23  
a. $\text{CS}_2(g) + 3\text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + 2\text{SO}_2(g)$  

b. 13.8 g of $\text{CO}_2$  
c. 37 mmHg  
d. 2370 mmHg

CI.25  
a. $K_c = 248$  
b. If $\text{H}_2$ is added, the equilibrium will shift in the direction of the products.  
c. If the volume decreases, the equilibrium will shift in the direction of the products.  
d. $[S_2] = 0.280 \text{ M}$

CI.27  
a. $2\text{M}(s) + 6\text{HCl}(aq) \rightarrow 3\text{H}_2(g) + 2\text{MCl}_3(aq)$  
b. 233 mL of $\text{H}_2$  
c. $6.03 \times 10^{-3}$ mol of $\text{M}$  
d. 69.7 g/mol; gallium  
e. $2\text{Ga}(s) + 6\text{HCl}(aq) \rightarrow 3\text{H}_2(g) + 2\text{GaCl}_3(aq)$

CI.29  
a. 0.180 M  
b. $[\text{H}_3\text{O}^+] = 5.56 \times 10^{-14} \text{ M}; \text{pH} = 13.255$  
c. $2\text{KOH}(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{K}_2\text{SO}_4(aq)$  
d. 3.60 mL

CI.31  
a. 0.001 20 M  
b. 0.0240 M  
c. $4.0 \times 10^{-8} \text{ M}$  
d. 7.40  
e. $\text{HCO}_3^-(aq) + \text{H}_3\text{O}^+(aq) \rightarrow \text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l)$  
f. $\text{H}_2\text{CO}_3(aq) + \text{OH}^-(aq) \rightarrow \text{HCO}_3^-(aq) + \text{H}_2\text{O}(l)$
Kimberly’s teeth have become badly stained by food, coffee, and drinks that form a layer on top of the enamel of her teeth. Kimberly’s dentist, Jane, has removed some of this layer by scraping her teeth, then brushing Kimberly’s teeth with an abrasive toothpaste. After that, Jane begins the process of whitening Kimberly’s teeth. She explains to Kimberly that she uses a gel containing 35% hydrogen peroxide (\( \text{H}_2\text{O}_2 \)) that penetrates into the enamel of the tooth, where it undergoes a chemical reaction that whitens the teeth. The chemical reaction is referred to as an oxidation–reduction reaction where one chemical (hydrogen peroxide) is reduced and the other chemical (the coffee stain) is oxidized. During the oxidation, the coffee stains on the teeth become lighter or colorless, and therefore, the teeth are whiter.

**CAREER**

**Dentist**

Dentists are involved in assessing and maintaining the health of the teeth, gums, mouth, and bones of the jaw. They also examine X-rays to determine if there is disease or decay of the teeth and gums. Dentists remove decay and fill cavities, or they may extract teeth if necessary. Local anesthetics are often used for dental procedures. Dentists also educate patients on proper care of teeth and gums. Many dentists operate their own businesses with a staff that includes dental hygienists and a dental technician. A dentist may specialize in certain areas of dentistry. An orthodontist fits a patient’s teeth with wires to straighten teeth or correct a bite. A periodontist treats problems with soft tissues of the gums. Pediatric dentists treat the teeth of children.
15.1 Oxidation and Reduction

**LEARNING GOAL** Identify a reaction as an oxidation or a reduction. Assign and use oxidation numbers to identify elements that are oxidized or reduced.

Oxidation and reduction are types of chemical reactions that are prevalent in our everyday lives. When we use natural gas to heat our home, turn on a computer, start a car, or eat food, we are using energy from the processes of oxidation and reduction.

In all of these processes, electrons are transferred during oxidation and reduction. In an oxidation reaction, a reactant loses one or more electrons. We say the substance is oxidized. In a reduction reaction, a reactant gains one or more electrons. We say the substance is reduced. When a substance undergoes oxidation, the electrons it loses must be gained by another substance, which is reduced. In an oxidation–reduction reaction, processes of oxidation and reduction must occur together. Sometimes we shorten this to redox reactions.

Writing Oxidation and Reduction Reactions

Perhaps you have seen glasses with photochromic lenses that darken when exposed to UV light from the Sun. Once the wearer returns indoors, the lenses become clear again. The change from clear to dark and back to clear is the result of oxidation and reduction reactions. We can use the compounds in the photochromic lenses to illustrate some oxidation and reduction reactions.

To make the lenses photosensitive, silver chloride (AgCl) and copper(I) chloride (CuCl) are embedded in the lens material. Away from the Sun, visible light goes through the clear lens. However, when exposed to the Sun, there is an oxidation–reduction reaction when the UV light is absorbed by AgCl, and Cl atoms and Ag atoms are produced. It is the silver atoms that are black and darken the lenses. This redox reaction is written:

\[
\text{AgCl} \rightarrow \text{Ag} + \text{Cl}^- \quad \text{or} \quad \text{Ag}^+ + \text{Cl}^- \rightarrow \text{Ag} + \text{Cl}^- \quad \text{Oxidation–reduction reaction}
\]

The chloride (Cl\(^-\)) ions lose electrons; they are oxidized. The silver (Ag\(^+\)) ions gain electrons; they are reduced.

**Half-Reactions**

To identify the oxidation and the reduction reactions, we can write each of the reactants with its product as a half-reaction. Then we observe that the Cl\(^-\) ion is oxidized because it loses an electron and the Ag\(^+\) ion is reduced because it gains an electron. The half-reaction
is called a reduction because the addition of an electron (negative charge) reduces the positive charge on Ag⁺.

\[
\text{Ag}^+ + e^- \rightarrow \text{Ag} \quad \text{Reduction} \quad \text{Gain of electron} \quad \text{Lenses darken}
\]

\[
\text{Cl}^- \rightarrow \text{Cl} + e^- \quad \text{Oxidation} \quad \text{Loss of electron}
\]

To prevent the lenses from clearing immediately, some Cu⁺ ions are added to the lenses. These Cu⁺ ions convert the Cl atoms produced from the darkening process to Cl⁻ ions.

\[
\text{Cu}^+ + \text{Cl} \rightarrow \text{Cu}^{2+} + \text{Cl}^- \quad \text{Oxidation}
\]

\[
\text{Cu}^+ \rightarrow \text{Cu}^{2+} + e^- \quad \text{Reduction}
\]

When the wearer of the lenses moves out of the sunlight, the reactions take place in the opposite direction. The Cu²⁺ ions combine with the silver atoms to re-form Cu⁺ and Ag⁺, which causes the lens to become clear again.

\[
\text{Cu}^{2+} + \text{Ag} \rightarrow \text{Cu}^+ + \text{Ag}^+ \quad \text{Reduction}
\]

\[
\text{Cu}^{2+} + e^- \rightarrow \text{Cu}^+ \quad \text{Oxidation} \quad \text{Lenses clear}
\]

**SAMPLE PROBLEM 15.1 Identifying Oxidation and Reduction Reactions**

For each of the following, identify the reaction as an oxidation or a reduction:

a. \(\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}\)  
   b. \(\text{Fe} \rightarrow \text{Fe}^{3+} + 3e^-\)  
   c. \(\text{Cr}^{3+} \rightarrow \text{Cr}^{3+} + e^-\)  
   d. \(\text{Li}^+ + e^- \rightarrow \text{Li}\)

**TRY IT FIRST**

**SOLUTION**

a. Since \(\text{Cu}^{2+}\) gained electrons, this is a reduction.

b. Since \(\text{Fe}\) lost electrons, this is an oxidation.

c. Since \(\text{Cr}^{3+}\) lost an electron, this is an oxidation.

d. Since \(\text{Li}^+\) gained an electron, this is a reduction.

**STUDY CHECK 15.1**

Is each of the following reactions an oxidation or a reduction?

a. \(\text{Na} \rightarrow \text{Na}^+ + e^-\)

b. \(\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}\)

**ANSWER**

a. oxidation  
   b. reduction

**Oxidation Numbers**

In more complex oxidation–reduction reactions, the identification of the substances oxidized and reduced is not always obvious. To help identify the atoms or ions that are oxidized or reduced, we assign values called *oxidation numbers* (sometimes called *oxidation states*) to the elements of the reactants and products. It is important to recognize that oxidation numbers do not always represent actual charges, but they help us identify loss or gain of electrons.

**Rules for Assigning Oxidation Numbers**

The rules for assigning oxidation numbers to the atoms or ions in the reactants and products are given in **TABLE 15.1**.
**TABLE 15.1 Rules for Assigning Oxidation Numbers**

1. The sum of the oxidation numbers in a molecule is zero (0), or for a polyatomic ion is equal to its charge.
2. The oxidation number of an element (monatomic or diatomic) is zero (0).
3. The oxidation number of a monatomic ion is equal to its charge.
4. In compounds, the oxidation number of Group 1A (1) metals is +1, and that of Group 2A (2) metals is +2.
5. In compounds, the oxidation number of fluorine is −1. Other nonmetals in Group 7A (17) are −1 except when combined with oxygen or fluorine.
6. In compounds, the oxidation number of oxygen is −2 except in OF₂, where O is +1; in H₂O₂ and other peroxides, O is −1.
7. In compounds with nonmetals, the oxidation number of hydrogen is +1; in compounds with metals, the oxidation number of hydrogen is −1.

We can now look at how these rules are used to assign oxidation numbers. For each formula, the oxidation numbers are written below the symbols of the elements (see Table 15.2).

**TABLE 15.2 Examples of Using Rules to Assign Oxidation Numbers**

<table>
<thead>
<tr>
<th>Formula</th>
<th>Oxidation Numbers</th>
<th>Explanation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Br₂</td>
<td>Br₂₀</td>
<td>Each Br atom in diatomic bromine has an oxidation number of 0 (Rule 2).</td>
</tr>
<tr>
<td>Ba²⁺</td>
<td>Ba²⁺²⁺</td>
<td>The oxidation number of a monatomic ion is equal to its charge (Rule 3).</td>
</tr>
<tr>
<td>CO₂</td>
<td>CO₂⁺⁻²</td>
<td>In compounds, O has an oxidation number of −2 (Rule 6). Because CO₃ is neutral, the oxidation number of C is calculated as +4 (Rule 1).</td>
</tr>
<tr>
<td>Al₂O₃</td>
<td>Al₂O₃⁺⁻²</td>
<td>In compounds, the oxidation number of O is −2 (Rule 6). For Al₂O₃ (neutral), the oxidation number of Al is calculated as +3 (Rule 1).</td>
</tr>
<tr>
<td>HClO₃</td>
<td>HClO₃⁺⁻²</td>
<td>In compounds, the oxidation number of H is +1 (Rule 7), and O is −2 (Rule 6). For HClO₃ (neutral), the oxidation number of Cl is calculated as +5 (Rule 1).</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>SO₄²⁻⁺⁻²</td>
<td>The oxidation number of O is −2 (Rule 6). For SO₄²⁻ (−2 charge), the oxidation number of S is calculated as +6 (Rule 1).</td>
</tr>
<tr>
<td>CH₂O</td>
<td>CH₂O₀⁺⁻²</td>
<td>In compounds, the oxidation number of H is +1 (Rule 7), and O is −2 (Rule 6). For CH₂O (neutral), the oxidation number of C is calculated as 0 (Rule 1).</td>
</tr>
</tbody>
</table>
SAMPLE PROBLEM 15.2 Assigning Oxidation Numbers

Assign oxidation numbers to the elements in each of the following:

a. NCl\textsubscript{3}  

b. CO\textsubscript{3}\textsuperscript{2-}

TRY IT FIRST

SOLUTION

a. NCl\textsubscript{3}: The oxidation number of Cl is \(-1\) (Rule 5). For NCl\textsubscript{3} (neutral), the sum of the oxidation numbers of N and 3Cl must be equal to zero (Rule 1). Thus, the oxidation number of N is calculated as +3.

\[
\begin{align*}
N + 3Cl &= 0 \\
N + 3(-1) &= 0 \\
N &= +3
\end{align*}
\]

The oxidation numbers are written as NCl\textsubscript{3}:

\[
N^{+3}Cl^{-3}\]

b. CO\textsubscript{3}\textsuperscript{2-}: The oxidation number of O is \(-2\) (Rule 6). For CO\textsubscript{3}\textsuperscript{2-}, the sum of the oxidation numbers is equal to \(-2\) (Rule 1). The oxidation number of C is calculated as +4.

\[
\begin{align*}
C + 3O &= -2 \\
C + 3(-2) &= -2 \\
C &= +4
\end{align*}
\]

The oxidation numbers are written as CO\textsubscript{3}\textsuperscript{2-}:

\[
C^{+4}O^{-2}\]

STUDY CHECK 15.2

Assign oxidation numbers to the elements in each of the following:

a. H\textsubscript{3}PO\textsubscript{4}  

b. MnO\textsubscript{4}^{-}

ANSWER

a. Because H has an oxidation number of +1, and O is \(-2\), P will have an oxidation number of +5 to maintain a neutral charge.

\[
H_{3}PO_{4}\\n_{+1}^{+5}O_{-2}
\]

b. Because O has an oxidation number of \(-2\), Mn must be +7 to give an overall charge of \(-1\).

\[
MnO_{4}^{-}\\n_{+7}^{-2}
\]

Using Oxidation Numbers to Identify Oxidation–Reduction

Oxidation numbers can be used to identify the elements that are oxidized and the elements that are reduced in a reaction. In oxidation, the loss of electrons increases the oxidation number so that it is higher (more positive) in the product than in the reactant. In reduction, the gain of electrons decreases the oxidation number so that it is lower (more negative) in the product than in the reactant.

Reduction: oxidation number decreases

\[-7 \rightarrow -6 \rightarrow -5 \rightarrow -4 \rightarrow -3 \rightarrow -2 \rightarrow -1 \rightarrow 0 \rightarrow +1 \rightarrow +2 \rightarrow +3 \rightarrow +4 \rightarrow +5 \rightarrow +6 \rightarrow +7\]

Oxidation: oxidation number increases

An oxidation reaction occurs when the charge becomes more positive. A reduction reaction occurs when the charge becomes more negative.
SAMPLE PROBLEM 15.3 Using Oxidation Numbers to Determine Oxidation and Reduction

Identify the element that is oxidized and the element that is reduced in the following equation:

$$\text{CO}_2(g) + \text{H}_2(g) \rightarrow \text{CO}(g) + \text{H}_2\text{O}(g)$$

**TRY IT FIRST**

**SOLUTION**

**STEP 1** Assign oxidation numbers to each element. In H₂, the oxidation number of H is 0. In H₂O, the oxidation number of H is +1. In CO₂, CO, and H₂O, the oxidation number of O is −2. Using the oxidation number −2 for O, the oxidation number of C would be +4 in CO₂, and +2 in CO.

$$\begin{align*}
\text{CO}_2(g) + \text{H}_2(g) &\rightarrow \text{CO}(g) + \text{H}_2\text{O}(g) \\
+4 &-2 &0 &+2 &-2 &+1 &-2 \quad \text{Oxidation numbers}
\end{align*}$$

**STEP 2** Identify the increase in oxidation number as oxidation; the decrease as reduction. H is oxidized because its oxidation number increases from 0 to +1. C is reduced because its oxidation number decreases from +4 to +2.

**STUDY CHECK 15.3**

In the following equation, which reactant is oxidized and which is reduced?

$$2\text{Al}(s) + 3\text{Sn}^{2+}(aq) \rightarrow 2\text{Al}^{3+}(aq) + 3\text{Sn}(s)$$

**ANSWER**

The oxidation number of Al increases from 0 to +3; Al is oxidized. The oxidation number of Sn²⁺ decreases from +2 to 0; Sn²⁺ is reduced.

**Oxidizing and Reducing Agents**

We have seen that an oxidation reaction must always be accompanied by a reduction reaction. The substance that loses electrons is oxidized, and the substance that gains electrons is reduced. For example, Zn is oxidized to Zn²⁺ by losing 2 electrons and Cl₂ reduced to 2Cl⁻ by gaining 2 electrons.

$$\text{Zn}(s) + \text{Cl}_2(g) \rightarrow \text{ZnCl}_2(s)$$

In an oxidation–reduction reaction, the substance that is oxidized is called the reducing agent because it provides electrons for reduction. The substance that is reduced is called the oxidizing agent because it accepts electrons from oxidation. Because Zn is oxidized in this reaction, it is the reducing agent. In the same reaction Cl₂ is reduced, which makes it the oxidizing agent.

$$\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^- \quad \text{Zn is oxidized; Zn is the reducing agent.}$$

$$\text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^- \quad \text{Cl (in Cl}_2\text{) is reduced; Cl}_2\text{ is the oxidizing agent.}$$

The terms we have used to describe oxidation and reduction are listed in the margin on the left.
SAMPLE PROBLEM 15.4 Identifying Oxidizing and Reducing Agents

Identify the substance that is oxidized, the substance that is reduced, the oxidizing agent, and the reducing agent for the reaction of lead(II) oxide and carbon monoxide.

PbO(s) + CO(g) → Pb(s) + CO₂(g)

TRY IT FIRST

SOLUTION

STEP 1 Assign oxidation numbers to the components in the equation.

Using the oxidation number −2 for O, the oxidation number of Pb in PbO is +2, the oxidation number of C in CO is +2, and the C in CO₂ would have an oxidation number of +4. The element Pb has an oxidation number of 0.

<table>
<thead>
<tr>
<th>Oxidized; reducing agent</th>
<th>PbO(s) + CO(g) → Pb(s) + CO₂(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>+2 −2</td>
<td>0 +4 −2</td>
</tr>
</tbody>
</table>

Reduced; oxidizing agent

STEP 2 Identify the substance that is oxidized and the substance that is reduced. The C in CO is oxidized because the oxidation number of C increases from +2 to +4. The Pb in PbO is reduced because the oxidation number of Pb decreases from +2 to 0.

STEP 3 Identify the oxidized substance as the reducing agent and the reduced substance as the oxidizing agent. The compound CO is oxidized; it is the reducing agent. The compound PbO is reduced; it is the oxidizing agent.

STUDY CHECK 15.4

Identify the oxidizing agent and the reducing agent in the following reaction:

2Al(s) + 3CuO(s) → Al₂O₃(s) + 3Cu(s)

ANSWER

CuO is the oxidizing agent and Al is the reducing agent.

QUESTIONS AND PROBLEMS

15.1 Oxidation and Reduction

LEARNING GOAL Identify a reaction as an oxidation or a reduction. Assign and use oxidation numbers to identify elements that are oxidized or reduced.

15.1 Identify each of the following as an oxidation or a reduction reaction:

a. Al⁺³(aq) + 3 e⁻ → Al(s)
b. Ca(s) → Ca⁺²(aq) + 2 e⁻
c. Fe⁺³(aq) + e⁻ → Fe⁺²(aq)

15.2 Identify each of the following as an oxidation or a reduction reaction:

a. Ni⁺²(aq) → Ni⁺⁺(aq) + e⁻
b. K⁺(aq) + e⁻ → K(s)
c. Br₂(l) + 2 e⁻ → 2Br⁻(aq)

15.3 In each of the following reactions, identify the reactant that is oxidized and the reactant that is reduced:

a. 4Al(s) + 3O₂(g) → 2Al₂O₃(s)
b. Zn(s) + 2H⁺(aq) → Zn⁺²(aq) + H₂(g)
c. F₂(g) + 2Br⁻(aq) → 2F⁻(aq) + Br₂(l)
15.4 In each of the following reactions, identify the reactant that is oxidized and the reactant that is reduced:
   a. $2\text{Ag}^+(aq) + \text{Zn}(s) \rightarrow 2\text{Ag}(s) + \text{Zn}^2+(aq)$
   b. $\text{Ca}(s) + \text{S}(s) \rightarrow \text{CaS}(s)$
   c. $\text{Sn}^2+(aq) + 2\text{Cr}^{3+}(aq) \rightarrow \text{Sn}^{4+}(aq) + 2\text{Cr}^{2+}(aq)$

15.5 Assign oxidation numbers to each of the following:
   a. Cu   b. F$_2$   c. Fe$^{2+}$   d. Cl$^-$

15.6 Assign oxidation numbers to each of the following:
   a. Al   b. Al$^{3+}$   c. I$^-$   d. N$_2$

15.7 Assign oxidation numbers to all the elements in each of the following:
   a. KCl   b. MnO$_2$   c. NO   d. Mn$_3$O$_4$

15.8 Assign oxidation numbers to all the elements in each of the following:
   a. H$_2$S   b. NO$_2$   c. CCl$_4$   d. PCl$_3$

15.9 Assign oxidation numbers to all the elements in each of the following:
   a. Li$_3$PO$_4$   b. SO$_3^{2-}$   c. Cr$_2$S$_3$   d. NO$_3^-$

15.10 Assign oxidation numbers to all the elements in each of the following:
   a. C$_2$H$_2$O$_2^-$   b. AlCl$_3$   c. NH$_4^+$   d. HBrO$_4$

15.11 Assign oxidation numbers to all the elements in each of the following:
   a. HSO$_4^-$   b. H$_2$PO$_3$   c. Cr$_2$O$_7^{2-}$   d. Na$_2$CO$_3$

15.12 Assign oxidation numbers to all the elements in each of the following:
   a. N$_2$O   b. LiOH   c. SbO$_2^-$   d. IO$_4^-$

15.13 What is the oxidation number of the specified element in each of the following?
   a. N in HNO$_3$   b. C in C$_3$H$_6$
   c. P in K$_3$PO$_4$   d. Cr in CrO$_3$

15.14 What is the oxidation number of the specified element in each of the following?
   a. C in BaCO$_3$   b. Fe in FeBr$_3$
   c. Cl in ClF$_3$   d. S in S$_2$O$_3^{2-}$

15.15 Indicate whether each of the following describes the oxidizing agent or the reducing agent in an oxidation–reduction reaction:
   a. the substance that is oxidized
   b. the substance that gains electrons

15.16 Indicate whether each of the following describes the oxidizing agent or the reducing agent in an oxidation–reduction reaction:
   a. the substance that is reduced
   b. the substance that loses electrons

15.17 For each of the following reactions, identify the substance that is oxidized, the substance that is reduced, the oxidizing agent, and the reducing agent:
   a. $2\text{Li}(s) + \text{Cl}_2(g) \rightarrow 2\text{LiCl}(s)$
   b. $\text{Cl}_2(g) + 2\text{NaBr}(aq) \rightarrow 2\text{NaCl}(aq) + \text{Br}_2(l)$
   c. $2\text{Pb}(s) + \text{O}_2(g) \rightarrow 2\text{PbO}(s)$
   d. $\text{Al}(s) + 3\text{Ag}^+(aq) \rightarrow \text{Al}^{3+}(aq) + 3\text{Ag}(s)$

15.18 For each of the following reactions, identify the substance that is oxidized, the substance that is reduced, the oxidizing agent, and the reducing agent:
   a. $2\text{Li}(s) + \text{F}_2(g) \rightarrow 2\text{LiF}(s)$
   b. $\text{Cl}_2(g) + 2\text{KI}(aq) \rightarrow 2\text{KCl}(aq) + \text{I}_2(s)$
   c. $\text{Zn}(s) + \text{Ni}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Ni}(s)$
   d. $\text{Fe}(s) + \text{CuSO}_4(aq) \rightarrow \text{FeSO}_4(aq) + \text{Cu}(s)$

15.19 For each of the following reactions, identify the substance that is oxidized, the substance that is reduced, the oxidizing agent, and the reducing agent:
   a. $2\text{NiS}(s) + 3\text{O}_2(g) \rightarrow 2\text{NiO}(s) + 2\text{SO}_2(g)$
   b. $\text{Sn}^{2+}(aq) + 2\text{Fe}^{3+}(aq) \rightarrow \text{Sn}^{4+}(aq) + 2\text{Fe}^{2+}(aq)$
   c. $\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)$
   d. $2\text{Cr}_2\text{O}_7^{2-}(s) + 3\text{Si}(s) \rightarrow 4\text{Cr}(s) + 3\text{SiO}_2(s)$

15.20 For each of the following reactions, identify the substance that is oxidized, the substance that is reduced, the oxidizing agent, and the reducing agent:
   a. $2\text{HgO}(s) \rightarrow 2\text{Hg}(l) + \text{O}_2(g)$
   b. $\text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)$
   c. $2\text{Na}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{Na}^{+}(aq) + 2\text{OH}^-(aq) + \text{H}_2(g)$
   d. $14\text{H}^+(aq) + 6\text{Fe}^{2+}(aq) + \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow 6\text{Fe}^{3+}(aq) + 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O}(l)$

15.2 Balancing Oxidation–Reduction Equations Using Half-Reactions

**LEARNING GOAL** Balance oxidation–reduction equations using the half-reaction method.

In the **half-reaction method** for balancing equations, an oxidation–reduction reaction is written as two **half-reactions**. The half-reaction method uses ionic charges and electrons to balance each half-reaction. Oxidation numbers are not used. The loss of electrons by one reactant is used to identify the oxidized substance and the gain of electrons by another reactant is used to identify the reduced substance. Once the loss and gain of electrons are equalized for the half-reactions, they are combined to obtain the overall balanced equation. The half-reaction method is typically used to balance equations that are written as ionic equations. Let us consider the reaction between aluminum metal and a solution of Cu$^{2+}$ as shown in Sample Problem 15.5.
SAMPLE PROBLEM 15.5  Using Half-Reactions to Balance Equations

Use half-reactions to balance the following equation:

\[ \text{Al}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Al}^{3+}(aq) + \text{Cu}(s) \]

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>Equation: Al(s) + Cu^{2+}(aq) \rightarrow Al^{3+}(aq) + Cu(s)</td>
<td>balanced equation</td>
<td>half-reactions, electron balance</td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Write two half-reactions for the equation.

\[
\begin{align*}
\text{Al}(s) & \rightarrow \text{Al}^{3+}(aq) \\
\text{Cu}^{2+}(aq) & \rightarrow \text{Cu}(s)
\end{align*}
\]

**STEP 2** For each half-reaction, balance the elements other than H and O. In these half-reactions, the Al and Cu are already balanced.

\[
\begin{align*}
\text{Al}(s) & \rightarrow \text{Al}^{3+}(aq) \\
\text{Cu}^{2+}(aq) & \rightarrow \text{Cu}(s)
\end{align*}
\]

**STEP 3** Balance each half-reaction for charge by adding electrons. For the aluminum half-reaction, we need to add three electrons on the product side to balance the charge. With a loss of electrons, this is an oxidation.

\[
\begin{align*}
\text{Al}(s) & \rightarrow \text{Al}^{3+}(aq) + 3e^- \\
0 \text{ charge} & = 0 \text{ charge}
\end{align*}
\]

For the Cu^{2+} half-reaction, we need to add two electrons on the reactant side to balance the charge. With a gain of electrons, this is a reduction.

\[
\begin{align*}
\text{Cu}^{2+}(aq) + 2e^- & \rightarrow \text{Cu}(s) \\
0 \text{ charge} & = 0 \text{ charge}
\end{align*}
\]

**STEP 4** Multiply each half-reaction by factors that equalize the loss and gain of electrons. To obtain the same number of electrons in each half-reaction, we need to multiply the oxidation half-reaction by 2 and the reduction half-reaction by 3.

\[
\begin{align*}
2 \times [\text{Al}(s) & \rightarrow \text{Al}^{3+}(aq) + 3e^-] \\
2\text{Al}(s) & \rightarrow 2\text{Al}^{3+}(aq) + 6e^- & 6e^- \text{ lost}
\end{align*}
\]

\[
\begin{align*}
3 \times [\text{Cu}^{2+}(aq) + 2e^- & \rightarrow \text{Cu}(s)] \\
3\text{Cu}^{2+}(aq) + 6e^- & \rightarrow 3\text{Cu}(s) & 6e^- \text{ gained}
\end{align*}
\]

**STEP 5** Add half-reactions and cancel electrons, and any identical ions or molecules. Check the balance of atoms and charge.

\[
\begin{align*}
2\text{Al}(s) & \rightarrow 2\text{Al}^{3+}(aq) + 6e^- \\
3\text{Cu}^{2+}(aq) + 6e^- & \rightarrow 3\text{Cu}(s)
\end{align*}
\]

\[
2\text{Al}(s) + 3\text{Cu}^{2+}(aq) + 6e^- \rightarrow 2\text{Al}^{3+}(aq) + 6e^- + 3\text{Cu}(s)
\]

**Final balanced equation:**

\[ 2\text{Al}(s) + 3\text{Cu}^{2+}(aq) \rightarrow 2\text{Al}^{3+}(aq) + 3\text{Cu}(s) \]
Check the balance of atoms and charge.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>2Al</td>
<td>2Al</td>
</tr>
<tr>
<td>3Cu</td>
<td>3Cu</td>
</tr>
<tr>
<td>Charge:</td>
<td>6+</td>
</tr>
</tbody>
</table>

**STUDY CHECK 15.5**
Use the half-reaction method to balance the following equation:

\[ \text{Zn}(s) + \text{Fe}^{3+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Fe}^{2+}(aq) \]

**ANSWER**

\[ \text{Zn}(s) + 2\text{Fe}^{3+}(aq) \rightarrow \text{Zn}^{2+}(aq) + 2\text{Fe}^{2+}(aq) \]

---

**Balancing Oxidation–Reduction Equations in Acidic Solution**

When we use the half-reaction method for balancing equations for reactions in acidic solution, we include balancing O by adding H₂O and balancing H by adding H⁺ as shown in Sample Problem 15.6.

**SAMPLE PROBLEM 15.6 Using Half-Reactions to Balance Equations in Acidic Solution**

Use half-reactions to balance the following equation for a reaction that takes place in acidic solution:

\[ \text{I}^- (aq) + \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow \text{I}_2(s) + \text{Cr}^{3+}(aq) \]

**TRY IT FIRST**

**SOLUTION**

**STEP 1** Write two half-reactions for the equation.

\[ \text{I}^- (aq) \rightarrow \text{I}_2(s) \]

\[ \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow \text{Cr}^{3+}(aq) \]

**STEP 2** For each half-reaction, balance the elements other than H and O. If the reaction takes place in acid, balance O by adding H₂O, and H by adding H⁺.

The two I atoms in I₂ are balanced with a coefficient of 2 for I⁻.

\[ 2\text{I}^- (aq) \rightarrow \text{I}_2(s) \]

The two Cr atoms are balanced with a coefficient of 2 for Cr³⁺.

\[ \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow 2\text{Cr}^{3+}(aq) \]

Now balance O by adding H₂O to the product side.

\[ \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O}(l) \quad \text{H}_2\text{O} \text{ balances O} \]

Balance H by adding H⁺ to the reactant side.

\[ 14\text{H}^+(aq) + \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O}(l) \quad \text{H}^+ \text{ balances H} \]

**STEP 3** Balance each half-reaction for charge by adding electrons. A charge of −2 is balanced with two electrons on the product side.

\[ 2\text{I}^- (aq) \rightarrow \text{I}_2(s) + 2e^- \quad \text{Oxidation} \]

A charge of +6 is obtained on the reactant side by adding six electrons.

\[ 6e^- + 14\text{H}^+(aq) + \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O}(l) \quad \text{Reduction} \]
Multiply each half-reaction by factors that equalize the loss and gain of electrons. The half-reaction with I is multiplied by 3 to equal the gain of $6\text{e}^-$ by the Cr half-reaction.

$$3 \times [2\text{I}^-(aq) \rightarrow \text{I}_2(s) + 2\text{e}^-]$$

$6\text{e}^- + 14\text{H}^+(aq) + \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O}(l)\quad 6\text{e}^- \text{ gained}$

Add half-reactions and cancel electrons, and any identical ions or molecules. Check the balance of atoms and charge.

$$6\text{I}^-(aq) \rightarrow 3\text{I}_2(s) + 6\text{e}^-$$

$6\text{e}^- + 14\text{H}^+(aq) + \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O}(l)$

Final balanced equation:

$$14\text{H}^+(aq) + \text{Cr}_2\text{O}_7^{2-}(aq) + 6\text{I}^-(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 3\text{I}_2(s) + 7\text{H}_2\text{O}(l)$$

Check the balance of atoms and charge.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>6I</td>
<td>6I</td>
</tr>
<tr>
<td>2Cr</td>
<td>2Cr</td>
</tr>
<tr>
<td>14H</td>
<td>14H</td>
</tr>
<tr>
<td>7O</td>
<td>7O</td>
</tr>
</tbody>
</table>

Charge: $6+ = 6+$

STUDY CHECK 15.6

Use half-reactions to balance the following equation in acidic solution:

$$\text{Fe}^{2+}(aq) + \text{IO}_3^-(aq) \rightarrow \text{Fe}^{3+}(aq) + \text{I}_2(s)$$

**ANSWER**

$$12\text{H}^+(aq) + 10\text{Fe}^{2+}(aq) + 2\text{IO}_3^-(aq) \rightarrow 10\text{Fe}^{3+}(aq) + \text{I}_2(s) + 6\text{H}_2\text{O}(l)$$

Balancing Oxidation–Reduction Equations in Basic Solution

An oxidation–reduction reaction can also take place in basic solution. In that case, we use the same half-reaction method, but once we have the balanced equation, we will neutralize the $\text{H}^+$ with $\text{OH}^-$ to form water. The $\text{H}^+$ is neutralized by adding $\text{OH}^-$ to both sides of the equation to form $\text{H}_2\text{O}$ as shown in Sample Problem 15.7.

**SAMPLE PROBLEM 15.7 Using Half-Reactions to Balance Equations in Basic Solution**

Use half-reactions to balance the following equation that takes place in basic solution:

$$\text{Fe}^{2+}(aq) + \text{MnO}_4^-(aq) \rightarrow \text{Fe}^{3+}(aq) + \text{MnO}_2(s)$$

**TRY IT FIRST**

**SOLUTION**

**STEP 1** Write two half-reactions for the equation.

$$\text{Fe}^{2+}(aq) \rightarrow \text{Fe}^{3+}(aq)$$

$$\text{MnO}_4^-(aq) \rightarrow \text{MnO}_2(s)$$
STEP 2 For each half-reaction, balance the elements other than H and O. If the reaction takes place in base, balance O by adding H₂O, and H by adding H⁺.

\[
\begin{align*}
\text{Fe}^{2+}(aq) & \rightarrow \text{Fe}^{3+}(aq) \\
\text{MnO}_4^-(aq) & \rightarrow \text{MnO}_2(s) + 2\text{H}_2\text{O}(l) & \text{H}_2\text{O} \text{ balances O} \\
4\text{H}^+(aq) + \text{MnO}_4^-(aq) & \rightarrow \text{MnO}_2(s) + 2\text{H}_2\text{O}(l) & \text{H}^+ \text{ balances H}
\end{align*}
\]

STEP 3 Balance each half-reaction for charge by adding electrons.

\[
\begin{align*}
\text{Fe}^{2+}(aq) & \rightarrow \text{Fe}^{3+}(aq) + 1e^- & \text{Oxidation} \\
+2 & = +2 \\
3e^- + 4\text{H}^+(aq) + \text{MnO}_4^-(aq) & \rightarrow \text{MnO}_2(s) + 2\text{H}_2\text{O}(l) & \text{Reduction}
\end{align*}
\]

STEP 4 Multiply each half-reaction by factors that equalize the loss and gain of electrons. The half-reaction with Fe is multiplied by 3 to equal the gain of 3 e⁻ by Mn.

\[
\begin{align*}
3 \times \text{[Fe}^{2+}(aq) & \rightarrow \text{Fe}^{3+}(aq) + 1e^-] \\
3\text{Fe}^{2+}(aq) & \rightarrow 3\text{Fe}^{3+}(aq) + 3e^- & 3e^- \text{ lost} \\
3e^- + 4\text{H}^+(aq) + \text{MnO}_4^-(aq) & \rightarrow \text{MnO}_2(s) + 2\text{H}_2\text{O}(l) & 3e^- \text{ gained}
\end{align*}
\]

STEP 5 Add half-reactions and cancel electrons, and any identical ions or molecules. In base, add OH⁻ to neutralize H⁺. Check the balance of atoms and charge.

\[
\begin{align*}
3\text{Fe}^{2+}(aq) & \rightarrow 3\text{Fe}^{3+}(aq) + 3e^- \\
3e^- + 4\text{H}^+(aq) + \text{MnO}_4^-(aq) & \rightarrow \text{MnO}_2(s) + 2\text{H}_2\text{O}(l)
\end{align*}
\]

\[
\begin{align*}
3e^- + 4\text{H}^+(aq) + 3\text{Fe}^{2+}(aq) + \text{MnO}_4^-(aq) & \rightarrow 3\text{Fe}^{3+}(aq) + \text{MnO}_2(s) + 2\text{H}_2\text{O}(l) + 3e^-
\end{align*}
\]

Final balanced equation:

\[
4\text{H}^+(aq) + 3\text{Fe}^{2+}(aq) + \text{MnO}_4^-(aq) \rightarrow 3\text{Fe}^{3+}(aq) + \text{MnO}_2(s) + 2\text{H}_2\text{O}(l)
\]

To convert the equation to an oxidation-reduction reaction in basic solution, we neutralize H⁺ with OH⁻ to form H₂O. For this equation, we add 4OH⁻(aq) to both sides of the equation.

\[
4\text{OH}^-(aq) + 4\text{H}^+(aq) + 3\text{Fe}^{2+}(aq) + \text{MnO}_4^-(aq) \rightarrow 3\text{Fe}^{3+}(aq) + \text{MnO}_2(s) + 2\text{H}_2\text{O}(l) + 4\text{OH}^-(aq)
\]

Combining 4H⁺ and 4OH⁻ gives 4H₂O on the reactant side.

\[
4\text{H}_2\text{O}(l) + 3\text{Fe}^{2+}(aq) + \text{MnO}_4^-(aq) \rightarrow 3\text{Fe}^{3+}(aq) + \text{MnO}_2(s) + 2\text{H}_2\text{O}(l) + 4\text{OH}^-(aq)
\]

Canceling 2H₂O on both the reactant and the product side gives the balanced equation in basic solution.

Final balanced equation in basic solution:

\[
2\text{H}_2\text{O}(l) + 3\text{Fe}^{2+}(aq) + \text{MnO}_4^-(aq) \rightarrow 3\text{Fe}^{3+}(aq) + \text{MnO}_2(s) + 4\text{OH}^-(aq)
\]

Check the balance of atoms and charge.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>3Fe</td>
<td>3Fe</td>
</tr>
<tr>
<td>1Mn</td>
<td>1Mn</td>
</tr>
<tr>
<td>4H</td>
<td>4H</td>
</tr>
<tr>
<td>6O</td>
<td>6O</td>
</tr>
<tr>
<td><strong>Charge:</strong> 5+</td>
<td><strong>5+</strong></td>
</tr>
</tbody>
</table>
15.3 Electrical Energy from Oxidation–Reduction Reactions

**LEARNING GOAL** Use the activity series to determine if an oxidation–reduction reaction is spontaneous. Write the half-reactions that occur in a voltaic cell and the cell notation.

When we placed a zinc metal strip in a solution of Cu^{2+}, reddish-brown Cu metal accumulated on the Zn strip according to the following spontaneous reaction.

\[ \text{Zn}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s) \quad \text{Spontaneous} \]

However, if we place a Cu metal strip in a Zn^{2+} solution, nothing will happen. The reaction does not run spontaneously in the reverse direction because Cu does not lose electrons as easily as Zn.

We can determine the direction of a spontaneous reaction from the **activity series**, which ranks the metals and H_2 in terms of how easily they lose electrons.

In the **activity series**, the metals that lose electrons most easily are placed at the top, and the metals that do not lose electrons easily are at the bottom. Thus the metals that are more easily oxidized are above the metals whose ions are more easily reduced (see **TABLE 15.3**). Active metals include K, Na, Ca, Mg, Al, Zn, Fe, and Sn. In single replacement reactions, the metal ion replaces the H in the acid. Metals listed below H_2(g) will not react with H^+ from acids.
TABLE 15.3 Activity Series for Some Metals

<table>
<thead>
<tr>
<th>Metal</th>
<th>Ion</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Most active</strong>&lt;br&gt;oxidize easily</td>
<td>Li(s) → Li⁺(aq) + e⁻</td>
</tr>
<tr>
<td>K(s) → K⁺(aq) + e⁻</td>
<td></td>
</tr>
<tr>
<td>Ca(s) → Ca²⁺(aq) + 2 e⁻</td>
<td></td>
</tr>
<tr>
<td>Na(s) → Na⁺(aq) + e⁻</td>
<td></td>
</tr>
<tr>
<td>Mg(s) → Mg²⁺(aq) + 2 e⁻</td>
<td></td>
</tr>
<tr>
<td>Al(s) → Al³⁺(aq) + 3 e⁻</td>
<td></td>
</tr>
<tr>
<td>Zn(s) → Zn²⁺(aq) + 2 e⁻</td>
<td></td>
</tr>
<tr>
<td>Cr(s) → Cr³⁺(aq) + 3 e⁻</td>
<td></td>
</tr>
<tr>
<td>Fe(s) → Fe²⁺(aq) + 2 e⁻</td>
<td></td>
</tr>
<tr>
<td>Ni(s) → Ni²⁺(aq) + 2 e⁻</td>
<td></td>
</tr>
<tr>
<td>Sn(s) → Sn²⁺(aq) + 2 e⁻</td>
<td></td>
</tr>
<tr>
<td>Pb(s) → Pb²⁺(aq) + 2 e⁻</td>
<td></td>
</tr>
<tr>
<td>H₂(g) → 2H⁺(aq) + 2 e⁻</td>
<td></td>
</tr>
<tr>
<td>Cu(s) → Cu²⁺(aq) + 2 e⁻</td>
<td></td>
</tr>
<tr>
<td>Ag(s) → Ag⁺(aq) + e⁻</td>
<td></td>
</tr>
<tr>
<td><strong>Least active</strong>&lt;br&gt;oxidize with difficulty</td>
<td>Au(s) → Au³⁺(aq) + 3 e⁻</td>
</tr>
</tbody>
</table>

According to the activity series, a metal will oxidize spontaneously when it is combined with the reverse of the half-reaction for any metal below it on the list. We use the activity series to predict the direction of the spontaneous reaction. Suppose we have two beakers. In one, we place a Mg strip in a solution containing Ni²⁺ ions. In the other, we place a Ni strip in a solution containing Mg²⁺ ions. Looking at the activity series we see that the half-reaction for the oxidation of Mg is listed above that for Ni, which means that Mg is the more active metal and loses electrons more easily than Ni. Using the activity series table, we write these two half-reactions as follows:

\[
\begin{align*}
\text{Mg(s)} & \rightarrow \text{Mg}^{2+}(aq) + 2 \text{e}^- \\
\text{Ni(s)} & \rightarrow \text{Ni}^{2+}(aq) + 2 \text{e}^-
\end{align*}
\]

The reaction that will be spontaneous is the oxidation of Mg combined with the reverse (reduction) of Ni²⁺.

\[
\begin{align*}
\text{Mg(s)} & \rightarrow \text{Mg}^{2+}(aq) + 2 \text{e}^- \\
\text{Ni}^{2+}(aq) + 2 \text{e}^- & \rightarrow \text{Ni(s)}
\end{align*}
\]

Therefore, we combine the half-reactions, which gives the following overall reaction that occurs spontaneously:

\[
\begin{align*}
\text{Mg(s)} + \text{Ni}^{2+}(aq) + 2 \text{e}^- & \rightarrow \text{Mg}^{2+}(aq) + \text{Ni(s)} + 2 \text{e}^- \\
\text{Mg(s)} + \text{Ni}^{2+}(aq) & \rightarrow \text{Mg}^{2+}(aq) + \text{Ni(s)}
\end{align*}
\]

However, a reaction between a Mg strip and a solution containing K⁺ ions will not occur spontaneously. We determine this by looking at the two half-reactions needed.

\[
\begin{align*}
\text{Mg(s)} & \rightarrow \text{Mg}^{2+}(aq) + 2 \text{e}^- \\
\text{K}^+(aq) + \text{e}^- & \rightarrow \text{K(s)}
\end{align*}
\]

Because the half-reaction for the oxidation of K is above that for Mg, there will not be a spontaneous reaction between Mg and K⁺.

\[
\text{Mg(s)} + 2\text{K}^+(aq) \cancel{ \rightarrow } \text{Mg}^{2+}(aq) + 2\text{K(s)}
\]

No reaction takes place
SAMPLE PROBLEM 15.8  Predicting Spontaneous Reactions

Determine if the reaction for each of the following metals with an HCl (H⁺) solution is spontaneous:

a. Zn(s) + 2H⁺(aq) → Zn²⁺(aq) + H₂(g)
b. Cu(s) + 2H⁺(aq) → Cu²⁺(aq) + H₂(g)

TRY IT FIRST

SOLUTION

a. Using the activity series (see Table 15.3), we see that Zn oxidizes more easily than H₂. Thus the oxidation half-reaction for Zn is combined with the reverse half-reaction for H₂.

$$\text{Zn(s)} \rightarrow \text{Zn}^{2+}(aq) + 2\text{e}^-$$

$$2\text{H}^+(aq) + 2\text{e}^- \rightarrow \text{H}_2(g)$$

$$\text{Zn(s)} + 2\text{H}^+(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{H}_2(g)$$  Spontaneous

b. Using the activity series (see Table 15.3), we see that H₂ oxidizes more easily than Cu. In this reaction, we would be combining an oxidation half-reaction with one that is above it in the activity series.

$$2\text{H}^+(aq) + 2\text{e}^- \rightarrow \text{H}_2(g)$$

$$\text{Cu(s)} \rightarrow \text{Cu}^{2+}(aq) + 2\text{e}^-$$

$$\text{Cu(s)} + 2\text{H}^+(aq) \nrightarrow \text{Cu}^{2+}(aq) + \text{H}_2(g)$$  Not spontaneous

STUDY CHECK 15.8

Determine if the following reaction is spontaneous:

2Al(s) + 3Cu²⁺(aq) → 2Al³⁺(aq) + 3Cu(s)

ANSWER

Yes, this reaction is spontaneous.

Voltaic Cells

We can generate electrical energy from a spontaneous oxidation–reduction reaction by using an apparatus called a voltaic cell. The two half-reactions still take place, but in a cell the electrons must flow through an external circuit. For example, a piece of zinc metal in a Cu²⁺ solution becomes coated with a rusty-brown coating of Cu, while the blue color (Cu²⁺) of the solution fades. The oxidation of the zinc metal provides electrons for the reduction of the Cu²⁺ ions. We can write the two half-reactions as

$$\text{Zn(s)} \rightarrow \text{Zn}^{2+}(aq) + 2\text{e}^- \quad \text{Oxidation}$$

$$\text{Cu}^{2+}(aq) + 2\text{e}^- \rightarrow \text{Cu(s)} \quad \text{Reduction}$$

The overall reaction is

$$\text{Zn(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu(s)}$$

As long as the Zn metal and Cu²⁺ ions are in the same container, the electrons are transferred directly from Zn to Cu²⁺. However, when the components of the two half-reactions are placed in separate containers, called half-cells, the electrons flow from one half-cell to the other, producing an electrical current. In each half-cell, there is a strip of metal, called an electrode, in contact with the ionic solution. The electrode where oxidation takes place is called the anode; the cathode is the electrode where reduction takes place. In this example, the anode is a zinc metal strip placed in a Zn²⁺(ZnSO₄) solution. The cathode is a copper metal strip placed in a Cu²⁺(CuSO₄) solution. In this voltaic cell, the Zn anode and Cu cathode are connected by a wire that allows electrons to move from the oxidation half-cell to the reduction half-cell.
**Anode is where** oxidation takes place; electrons are produced:

\[ Zn(s) \rightarrow Zn^{2+}(aq) + 2 \, e^- \]

**Cathode is where** reduction takes place; electrons are used up:

\[ Cu^{2+}(aq) + 2 \, e^- \rightarrow Cu(s) \]

The circuit is completed by a *salt bridge* containing positive and negative ions that are placed in the half-cell solutions. The purpose of the salt bridge is to provide ions, such as \( \text{Na}^+ \) and \( \text{SO}_4^{2-} \) ions, to maintain an electrical balance in each half-cell solution. As oxidation occurs, there is an increase in \( Zn^{2+} \) ions, which is balanced by \( \text{SO}_4^{2-} \) anions from the salt bridge. At the cathode, there is a loss of \( Cu^{2+} \), which is balanced by \( \text{SO}_4^{2-} \) moving into the salt bridge. The complete circuit involves the flow of electrons from the anode to the cathode and the flow of anions from the cathode solution to the anode solution (see **FIGURE 15.1**).

**FIGURE 15.1** In this voltaic cell, the Zn anode is in a \( Zn^{2+} \) solution, and the Cu cathode is in a \( Cu^{2+} \) solution. Electrons produced by the oxidation of Zn flow from the anode through the wire to the cathode where they reduce \( Cu^{2+} \) to Cu. The circuit is completed by the flow of \( \text{SO}_4^{2-} \) through the salt bridge.

*Which electrode will be heavier when the reaction ends?*

We can diagram the oxidation and reduction reactions that take place in the cell using a *shorthand notation* as follows:

\[ Zn(s) | Zn^{2+}(aq) \parallel Cu^{2+}(aq) | Cu(s) \]

The components of the oxidation half-cell (anode) are written on the left side, and the components of the reduction half-cell (cathode) are written on the right. A single vertical line separates the solid Zn anode from the \( Zn^{2+} \) solution, and another vertical line separates the \( Cu^{2+} \) solution from the Cu cathode. A double vertical line separates the two half-cells.
In some voltaic cells, there is no component in the half-reactions that can be used as
an electrode. When this is the case, inert electrodes made of graphite or platinum are used
for the transfer of electrons. If there are two ionic components in a cell, their symbols are
separated by a comma. For example, suppose a voltaic cell consists of a platinum anode
placed in a \( \text{Sn}^{2+} \) solution, and a silver cathode placed in a \( \text{Ag}^+ \) solution. The notation for
the cell would be written as

\[
\text{Pt(s)} | \text{Sn}^{2+}(aq), \text{Sn}^{4+}(aq) \parallel \text{Ag}^+(aq) | \text{Ag(s)}
\]

The oxidation reaction at the anode is

\[
\text{Sn}^{2+}(aq) \rightarrow \text{Sn}^{4+}(aq) + 2 \text{e}^-
\]

The reduction reaction at the cathode is

\[
\text{Ag}^+(aq) + \text{e}^- \rightarrow \text{Ag(s)}
\]

To balance the overall cell reaction, we multiply the cathode reduction by 2 and combine
the two half-reactions.

\[
\begin{align*}
2\text{Ag}^+(aq) + 2 \text{e}^- & \rightarrow 2\text{Ag(s)} \quad \text{Reduction} \\
\text{Sn}^{2+}(aq) & \rightarrow \text{Sn}^{4+}(aq) + 2 \text{e}^- \quad \text{Oxidation} \\
\text{Sn}^{2+}(aq) + 2\text{Ag}^+(aq) & \rightarrow \text{Sn}^{4+}(aq) + 2\text{Ag(s)} \quad \text{Oxidation–Reduction Reaction}
\end{align*}
\]

**SAMPLE PROBLEM 15.9 Diagramming a Voltaic Cell**

A voltaic cell consists of an iron (Fe) anode in a \( \text{Fe}^{2+} \) solution and a tin (Sn) cathode
placed in a \( \text{Sn}^{2+} \) solution. Write the cell notation, the oxidation and reduction half-
reactions, and the overall cell reaction.

**TRY IT FIRST**

**SOLUTION**

The notation for the cell would be written as

\[
\text{Fe(s)} | \text{Fe}^{2+}(aq) \parallel \text{Sn}^{2+}(aq) \parallel \text{Sn(s)}
\]

The oxidation reaction at the anode is

\[
\text{Fe(s)} \rightarrow \text{Fe}^{2+}(aq) + 2 \text{e}^-
\]

The reduction reaction at the cathode is

\[
\text{Sn}^{2+}(aq) + 2 \text{e}^- \rightarrow \text{Sn(s)}
\]

To write the overall cell reaction, we combine the two half-reactions.

\[
\begin{align*}
\text{Fe(s)} & \rightarrow \text{Fe}^{2+}(aq) + 2 \text{e}^- \quad \text{Oxidation} \\
\text{Sn}^{2+}(aq) + 2 \text{e}^- & \rightarrow \text{Sn(s)} \quad \text{Reduction} \\
\text{Fe(s)} + \text{Sn}^{2+}(aq) & \rightarrow \text{Fe}^{2+}(aq) + \text{Sn(s)} \quad \text{Oxidation–Reduction Reaction}
\end{align*}
\]

**STUDY CHECK 15.9**

Write the half-reactions and the overall cell reaction for the following notation of a
voltaic cell:

\[
\text{Co(s)} | \text{Co}^{2+}(aq) \parallel \text{Cu}^{2+}(aq) \parallel \text{Cu(s)}
\]

**ANSWER**

Anode reaction: \( \text{Co(s)} \rightarrow \text{Co}^{2+}(aq) + 2 \text{e}^- \)

Cathode reaction: \( \text{Cu}^{2+}(aq) + 2 \text{e}^- \rightarrow \text{Cu(s)} \)

Overall cell reaction: \( \text{Co(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Co}^{2+}(aq) + \text{Cu(s)} \)
Batteries are needed to power your cell phone, watch, and calculator. Batteries also are needed to make cars start, and flashlights produce light. Within each of these batteries are voltaic cells that produce electrical energy. Let’s look at some examples of commonly used batteries.

**Lead Storage Battery**

A lead storage battery is used to operate the electrical system in a car. We need a car battery to start the engine, turn on the lights, or operate the radio. If the battery runs down, the car won’t start and the lights won’t turn on. A car battery or a lead storage battery is a type of voltaic cell. In a typical 12-V battery, there are six voltaic cells linked together. Each of the cells consists of a lead (Pb) plate that acts as the anode and a lead(IV) oxide (PbO₂) plate that acts as the cathode. Both half-cells contain a sulfuric acid (H₂SO₄) solution. When the car battery is producing electrical energy (discharging), the following half-reactions take place:

**Anode (oxidation):**

\[ \text{Pb}(s) + \text{SO}_4^{2-}(aq) \rightarrow \text{PbSO}_4(s) + 2e^- \]

**Cathode (reduction):**

\[ 2e^- + 4\text{H}^+(aq) + \text{PbO}_2(s) + \text{SO}_4^{2-}(aq) \rightarrow \text{PbSO}_4(s) + 2\text{H}_2\text{O}(l) \]

Overall cell reaction:

\[ 4\text{H}^+(aq) + \text{Pb}(s) + \text{PbO}_2(s) + 2\text{SO}_4^{2-}(aq) \rightarrow 2\text{PbSO}_4(s) + 2\text{H}_2\text{O}(l) \]

In both half-reactions, Pb²⁺ is produced, which combines with SO₄²⁻ to form an insoluble ionic compound PbSO₄. As a car battery is used, there is a buildup of PbSO₄ on the electrodes. At the same time, there is a decrease in the concentrations of the sulfuric acid components, H⁺ and SO₄²⁻. As a car runs, the battery is continuously recharged by an alternator, which is powered by the engine. The recharging reactions restore the Pb and PbO₂ electrodes as well as H₂SO₄. Without recharging, the car battery cannot continue to produce electrical energy.

**Dry-Cell Batteries**

Dry-cell batteries are used in calculators, watches, flashlights, and battery-operated toys. The term dry cell describes a battery that uses a paste rather than an aqueous solution. Dry cells can be acidic or alkaline. In an acidic dry cell, the anode is a zinc metal case that contains a paste of MnO₂, NH₄Cl, ZnCl₂, H₂O, and starch. Within this MnO₂ electrolyte mixture is a graphite cathode.

![Acidic dry-cell battery](image)
Anode (oxidation): \[ \text{Zn(s)} \rightarrow \text{Zn}^{2+}(aq) + 2 \text{e}^- \] 
Cathode (reduction): \[ 2 \text{e}^- + 2\text{MnO}_2(s) + 2\text{NH}_3(aq) \rightarrow \text{Mn}_2\text{O}_3(s) + 2\text{NH}_3(aq) + \text{H}_2\text{O}(l) \] 
Overall cell reaction: \[ \text{Zn(s)} + 2\text{MnO}_2(s) + 2\text{NH}_3(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Mn}_2\text{O}_3(s) + 2\text{NH}_3(aq) + \text{H}_2\text{O}(l) \]

An alkaline battery has similar components except that NaOH or KOH replaces the NH₄Cl electrolyte. Under basic conditions, the product of oxidation is zinc oxide (ZnO). Alkaline batteries tend to be more expensive, but they last longer and produce more power than acidic dry-cell batteries.

Anode (oxidation): \[ \text{Zn(s)} + 2\text{OH}^-(aq) \rightarrow \text{ZnO}(s) + \text{H}_2\text{O}(l) + 2 \text{e}^- \] 
Cathode (reduction): \[ 2 \text{e}^- + 2\text{NiO(OH)}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{Ni(OH)}_2(s) + 2\text{OH}^-(aq) \] 
Overall cell reaction: \[ \text{Zn(s)} + 2\text{NiO(OH)}(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{ZnO}(s) + 2\text{Ni(OH)}_2(s) \]

Nickel–Cadmium (NiCad) Batteries
Nickel–cadmium (NiCad) batteries can be recharged. They use a cadmium anode and a cathode of solid nickel oxide NiO(OH)(s).

Anode (oxidation): \[ \text{Cd(s)} + 2\text{OH}^-(aq) \rightarrow \text{Cd(OH)}_2(s) + 2 \text{e}^- \] 
Cathode (reduction): \[ 2 \text{e}^- + 2\text{NiO(OH)}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{Ni(OH)}_2(s) + 2\text{OH}^-(aq) \] 
Overall cell reaction: \[ \text{Cd(s)} + 2\text{NiO(OH)}(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Cd(OH)}_2(s) + 2\text{Ni(OH)}_2(s) \]

NiCad batteries are expensive, but they can be recharged many times. A charger provides an electrical current that converts the solid Cd(OH)₂ and Ni(OH)₂ products in the NiCad battery back to the reactants.

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**CHEMISTRY LINK TO THE ENVIRONMENT**

**Corrosion: Oxidation of Metals**

Metals used in building materials, such as iron, eventually oxidize, which causes deterioration of the metal. This oxidation process, known as corrosion, produces rust on cars, bridges, ships, and underground pipes.

\[ 4\text{Fe(s)} + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \]

Rust

The formation of rust requires both oxygen and water. The process of rusting requires an anode and cathode in different places on the surface of a piece of iron. In one area of the iron surface, called the **anode region**, the oxidation half-reaction takes place (see **FIGURE 15.2**).

Anode (oxidation): \[ \text{Fe(s)} \rightarrow \text{Fe}^{2+}(aq) + 2 \text{e}^- \] 
or
\[ 2\text{Fe(s)} \rightarrow 2\text{Fe}^{2+}(aq) + 4 \text{e}^- \]

The electrons move through the iron metal from the anode to an area called the **cathode region** where oxygen dissolved in water is reduced to water.

Cathode (reduction): \[ 4 \text{e}^- + 4\text{H}^+(aq) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) \]

By combining the half-reactions that occur in the anode and cathode regions, we can write the overall oxidation–reduction equation.

\[ 2\text{Fe(s)} + 4\text{H}^+(aq) + \text{O}_2(g) \rightarrow 2\text{Fe}^{2+}(aq) + 2\text{H}_2\text{O}(l) \]

**FIGURE 15.2** Rust forms when electrons from the oxidation of Fe flow from the anode region to the cathode region where oxygen is reduced. As Fe²⁺ ions come in contact with O₂ and H₂O, rust forms.

Why must both O₂ and H₂O be present for the corrosion of iron?

The formation of rust occurs as Fe²⁺ ions move out of the anode region and come in contact with dissolved oxygen (O₂). The Fe²⁺ oxidizes to Fe³⁺, which reacts with oxygen to form rust.

\[ 4\text{H}_2\text{O}(l) + 4\text{Fe}^{2+}(aq) + \text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) + 8\text{H}^+(aq) \]

Rust
We can write the formation of rust starting with solid Fe reacting with O₂ as follows. There is no H⁺ in the overall equation because H⁺ is used and produced in equal quantities.

Corrosion of iron

\[ 4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \]

Rust

Other metals such as aluminum, copper, and silver also undergo corrosion, but at a slower rate than iron. The oxidation of Al on the surface of an aluminum object produces Al⁺³, which reacts with oxygen in the air to form a protective coating of Al₂O₃. This Al₂O₃ coating prevents further oxidation of the aluminum underneath it.

\[ \text{Al}(s) \rightarrow \text{Al}^{3+}(aq) + 3e^- \]

When copper is used on a roof, dome, or a steeple, it oxidizes to Cu⁺², which is converted to a green patina of Cu₂(OH)₂CO₃.

\[ \text{Cu}(s) \rightarrow \text{Cu}^{2+}(aq) + 2e^- \]

When we use silver dishes and utensils, the Ag⁺ ion from oxidation reacts with sulfides in food to form Ag₂S, which we call “tarnish.”

\[ \text{Ag}(s) \rightarrow \text{Ag}^{+}(aq) + e^- \]

Prevention of Corrosion

Billions of dollars are spent each year to prevent corrosion and repair building materials made of iron. One way to prevent corrosion is to paint the bridges, cars, and ships with paints containing materials that seal the iron surface from H₂O and O₂. But it is necessary to repaint often; a scratch in the paint exposes the iron, which then begins to rust.

A more effective way to prevent corrosion is to place the iron in contact with a metal that substitutes for the anode region of iron. Metals such as Zn, Mg, or Al lose electrons more easily than iron. When one of these metals is in contact with iron, the metal acts as the anode instead of iron. For example, in a process called galvanization, an iron object is coated with zinc. The zinc becomes the anode because zinc is more easily oxidized than Fe. As long as Fe does not act as an anode, rust cannot form.

In a method called cathodic protection, structures such as iron pipes and underground storage containers are placed in contact with a piece of metal such as Mg, Al, or Zn, which is called the sacrificial anode. Again, because these metals lose electrons more easily than Fe, they become the anode, thereby preventing the rusting of the iron. A magnesium plate that is welded or bolted to a ship’s hull loses electrons more easily than iron or steel and protects the hull from rusting. Occasionally, a new magnesium plate is added to replace the magnesium as it is used up. Magnesium stakes placed in the ground are connected to underground pipelines and storage containers to prevent corrosion damage.

**QUESTIONS AND PROBLEMS**

**15.3 Electrical Energy from Oxidation–Reduction Reactions**

**LEARNING GOAL** Use the activity series to determine if an oxidation–reduction reaction is spontaneous. Write the half-reactions that occur in a voltaic cell and the cell notation.

15.27 Use the activity series in Table 15.3 to predict whether each of the following reactions will occur spontaneously:

- a. 2Au(s) + 6H⁺(aq) → 2Au⁺³(aq) + 3H₂(g)
- b. Ni²⁺(aq) + Fe(s) → Ni(s) + Fe⁺²(aq)
- c. 2Ag(s) + Cu²⁺(aq) → 2Ag⁺(aq) + Cu(s)

15.28 Use the activity series in Table 15.3 to predict whether each of the following reactions will occur spontaneously:

- a. 2Ag(s) + 2H⁺(aq) → 2Ag⁺(aq) + H₂(g)
- b. Mg(s) + Cu²⁺(aq) → Mg²⁺(aq) + Cu(s)
- c. 2Al(s) + 3Cu²⁺(aq) → 2Al⁺³(aq) + 3Cu(s)

15.29 Write the half-reactions and the overall cell reaction for each of the following voltaic cells:

- a. Pb(s) | Pb⁺²(aq) || Cu⁺²(aq) | Cu(s)
- b. Cr(s) | Cr⁺³(aq) || Ag⁺(aq) | Ag(s)

15.30 Write the half-reactions and the overall cell reaction for each of the following voltaic cells:

- a. Al(s) | Al⁺³(aq) || Cd²⁺(aq) | Cd(s)
- b. Sn(s) | Sn⁺²(aq) || Fe³⁺(aq), Fe⁺²(aq) | C (electrode)

15.31 Describe the voltaic cell and half-cell components and write the shorthand notation for the following oxidation–reduction reactions:

- a. Cd(s) + Sn⁺²(aq) → Cd⁺²(aq) + Sn(s)
- b. Zn(s) + Cl⁺⁺(aq) → Zn⁺²(aq) + 2Cl⁻(aq) C (electrode)
15.32 Describe the voltaic cell and half-cell components and write the shorthand notation for the following oxidation–reduction reactions:

- a. \( \text{Mn}(s) + \text{Sn}^2(aq) \rightarrow \text{Mn}^2(aq) + \text{Sn}(s) \)
- b. \( \text{Ni}(s) + 2\text{Ag}^+(aq) \rightarrow \text{Ni}^2(aq) + 2\text{Ag}(s) \)

15.33 The following half-reaction takes place in a nickel–cadmium battery used in a cordless drill:

\[ \text{Cd}(s) + 2\text{OH}^-(aq) \rightarrow \text{Cd}^2+ + 2\text{OH}^- + 2e^- \]

- a. Is the half-reaction an oxidation or a reduction?
- b. What substance is oxidized or reduced?
- c. At which electrode would this half-reaction occur?

15.34 The following half-reaction takes place in a mercury battery used in hearing aids:

\[ \text{HgO}(s) + \text{H}_2\text{O}(l) + 2e^- \rightarrow \text{Hg}(l) + 2\text{OH}^-(aq) \]

- a. Is the half-reaction an oxidation or a reduction?
- b. What substance is oxidized or reduced?
- c. At which electrode would this half-reaction occur?

15.35 The following half-reaction takes place in a mercury battery used in pacemakers and watches:

\[ \text{Zn}(s) + 2\text{OH}^-(aq) \rightarrow \text{ZnO}(s) + \text{H}_2\text{O}(l) + 2e^- \]

- a. Is the half-reaction an oxidation or a reduction?
- b. What substance is oxidized or reduced?
- c. At which electrode would this half-reaction occur?

15.36 The following half-reaction takes place in a lead storage battery used in automobiles:

\[ \text{Pb}(s) + \text{SO}_4^{2-}(aq) \rightarrow \text{PbSO}_4(s) + 2e^- \]

- a. Is the half-reaction an oxidation or a reduction?
- b. What substance is oxidized or reduced?
- c. At which electrode would this half-reaction occur?

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**CHEMISTRY LINK TO THE ENVIRONMENT**

**Fuel Cells: Clean Energy for the Future**

Fuel cells are of interest to scientists because they provide an alternative source of electrical energy that is more efficient, does not use up oil reserves, and generates products that do not pollute the atmosphere. Fuel cells are considered to be a clean way to produce energy.

Like other electrochemical cells, a fuel cell consists of an anode and a cathode connected by a wire. But unlike other cells, the reactants must continuously enter the fuel cell to produce energy; electrical current is generated only as long as the fuels are supplied. One type of hydrogen–oxygen fuel cell has been used in automobile prototypes. In this hydrogen cell, gas enters the fuel cell and comes in contact with a platinum catalyst embedded in a plastic membrane. The catalyst assists in the oxidation of hydrogen atoms to hydrogen ions and electrons (see **FIGURE 15.3**).

The electrons produce an electric current as they travel through the wire from the anode to the cathode. The hydrogen ions flow through the plastic membrane to the cathode. At the cathode, oxygen molecules are reduced to oxide ions that combine with the hydrogen ions to form water. The overall hydrogen–oxygen fuel cell reaction can be written as

\[ 2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) \]

Fuel cells have already been used for power on the space shuttle and are now used to produce energy for cars and buses. An advantage of fuel cell cars is that the hydrogen fuel produces water and heat, which gives zero pollution.

**FIGURE 15.3** With a supply of hydrogen and oxygen, a fuel cell can generate electricity continuously.

(Question) In most electrochemical cells, the electrodes are eventually used up. Is this true for a fuel cell? Why or why not?
In homes, fuel cells may one day replace the batteries currently used to provide electrical power for cell phones, DVD players, and laptop computers. Fuel cell design is still in the prototype phase, although there is much interest in their development. We already know they can work, but modifications must still be made before they become reasonably priced and part of our everyday lives.

**15.4 Oxidation–Reduction Reactions That Require Electrical Energy**

**LEARNING GOAL** Describe the half-cell reactions and the overall reactions that occur in electrolysis.

When we look at the activity series in Table 15.3, we see that the oxidation of Cu is below Zn. This means that the following oxidation–reduction reaction is not spontaneous:

\[
\text{Cu}(s) + \text{Zn}^{2+}(aq) \rightarrow \text{Cu}^{2+}(aq) + \text{Zn}(s) \\
\text{Less active} \quad \text{Not spontaneous} \quad \text{More active}
\]

To make this reaction take place, we need an electrolytic cell that uses an electrical current to drive a nonspontaneous oxidation–reduction reaction. This process is called electrolysis (see Figure 15.4).

**FIGURE 15.4** In this electrolytic cell, the Cu anode is in a Cu\(^{2+}\) solution, and the Zn cathode is in a Zn\(^{2+}\) solution. Electrons provided by a battery reduce Zn\(^{2+}\) to Zn and drive the oxidation of Cu to Cu\(^{2+}\) at the Cu anode.

Why does the reaction in an electrolytic cell require an electrical current?

**Electrolysis of Sodium Chloride**

When molten sodium chloride is electrolyzed, the products are sodium metal and chlorine gas. In this electrolytic cell, electrodes are placed in the mixture of Na\(^+\) and Cl\(^-\) and connected to a battery. The products are separated to prevent them from reacting.
spontaneously with each other. As electrons flow to the cathode, \( \text{Na}^+ \) is reduced to sodium metal. At the same time, electrons leave the anode as \( \text{Cl}^- \) is oxidized to \( \text{Cl}_2 \). The half-reactions and the overall reactions are

\[
\begin{align*}
\text{Anode (oxidation):} & \quad 2\text{Cl}^- (l) & \rightarrow & \text{Cl}_2 (g) + 2e^- \\
\text{Cathode (reduction):} & \quad 2\text{Na}^+ (l) + 2e^- & \rightarrow & 2\text{Na}(l) \\
\text{Overall reaction:} & \quad 2\text{Na}^+ (l) + 2\text{Cl}^- (l) & \rightarrow & 2\text{Na}(l) + \text{Cl}_2 (g) \quad \text{Not spontaneous}
\end{align*}
\]

### Electroplating

In industry, the process of electroplating uses electrolysis to coat an object with a thin layer of a metal such as silver, platinum, or gold. Car bumpers are electroplated with chromium to prevent rusting. Silver-plated utensils, bowls, and platters are made by electroplating objects with a layer of silver.

#### SAMPLE PROBLEM 15.10  Electrolysis

Electrolysis is used to chrome plate an iron hubcap by placing the hubcap in a \( \text{Cr}^{3+} \) solution.

**a.** What half-reaction takes place to plate the hubcap with metallic chromium?

**b.** Is the iron hubcap the anode or the cathode?

**TRY IT FIRST**

**SOLUTION**

**a.** The \( \text{Cr}^{3+} \) ions in solution would gain electrons (reduction).

\[
\text{Cr}^{3+} (aq) + 3e^- \rightarrow \text{Cr}(s)
\]

**b.** The iron hubcap is the cathode where reduction takes place.

**STUDY CHECK 15.10**

Why is energy needed to chrome plate the iron in Sample Problem 15.10?

**ANSWER**

Since Fe is below Cr in the activity series, the plating of \( \text{Cr}^{3+} \) onto Fe is not spontaneous. Energy is needed to make the reaction proceed.
QUESTIONS AND PROBLEMS

15.4 Oxidation–Reduction Reactions That Require Electrical Energy

LEARNING GOAL Describe the half-cell reactions and the overall reactions that occur in electrolysis.

15.37 What we call “tin cans” are really iron cans coated with a thin layer of tin. The anode is a bar of tin and the cathode is the iron can. An electrical current is used to oxidize the Sn to Sn\(^{2+}\) in solution, which is reduced to produce a thin coating of Sn on the can.

a. What half-reaction takes place to tin plate an iron can?
b. Why is the iron can the cathode?
c. Why is the tin bar the anode?

15.38 Electrolysis is used to gold plate jewelry made of stainless steel.

a. What half-reaction takes place when a Au\(^{3+}\) solution is used to gold plate a stainless steel ring?
b. Is the ring the anode or the cathode?
c. Why is energy needed to gold plate the ring?

15.39 When the tin coating on an iron can is scratched, rust will form. Use the activity series in Table 15.3 to explain why this happens.

15.40 When the zinc coating on an iron can is scratched, rust does not form. Use the activity series in Table 15.3 to explain why this happens.

Follow Up

WHITENING KIMBERLY’S TEETH

The reaction of hydrogen peroxide, H\(_2\)O\(_2\)(aq), placed on Kimberly’s teeth produces water, H\(_2\)O(l), and oxygen gas. After Kimberly had her teeth cleaned and whitened, Jane recommended that Kimberly use a toothpaste containing tin(II) fluoride. Tooth enamel, which is composed of hydroxyapatite, Ca\(_5\)(PO\(_4\))\(_3\)OH, is strengthened when it reacts with fluoride ions to form fluoroapatite, Ca\(_5\)(PO\(_4\))\(_3\)F.

15.41 The simplified reaction for the production of aqueous H\(_2\)O\(_2\) involves the combination of H\(_2\) gas and O\(_2\) gas.

a. What reactant is oxidized and what reactant is reduced in this reaction?
b. What is the oxidizing agent?
c. What is the reducing agent?
d. Write a balanced chemical equation for the reaction.

15.42 The reaction of solid Sn and F\(_2\) gas forms solid SnF\(_2\).

a. What reactant is oxidized and what reactant is reduced in this reaction?
b. What is the oxidizing agent?
c. What is the reducing agent?
d. Write a balanced chemical equation for the reaction.


**CHAPTER REVIEW**

**15.1 Oxidation and Reduction**

**LEARNING GOAL** Identify a reaction as an oxidation or a reduction. Assign and use oxidation numbers to identify elements that are oxidized or reduced.

- In an oxidation–reduction reaction, electrons are transferred from one reactant to another.
- The reactant that loses electrons is oxidized, and the reactant that gains electrons is reduced.
- Oxidation numbers assigned to elements keep track of the changes in the loss and gain of electrons.
- Oxidation is an increase in oxidation number; reduction is a decrease in oxidation number.
- The reducing agent is the substance that provides electrons for reduction.
- The oxidizing agent is the substance that accepts the electrons from oxidation.
- In molecular compounds and polyatomic ions, oxidation numbers are assigned using a set of rules.
- The oxidation number of an element is zero, and the oxidation number of a monatomic ion is the same as the ionic charge of the ion.
- The sum of the oxidation numbers for a compound is equal to zero and for a polyatomic ion is equal to the overall charge.
- Balancing oxidation–reduction equations using oxidation numbers involves the following:
  1. assigning oxidation numbers
  2. determining the loss and gain of electrons
  3. equalizing the loss and gain of electrons
  4. balancing the remaining substances by inspection

**15.2 Balancing Oxidation–Reduction Equations Using Half-Reactions**

**LEARNING GOAL** Balance oxidation–reduction equations using the half-reaction method.

- Balancing oxidation–reduction equations using half-reactions involves the following:
  1. separating the equation into half-reactions
  2. balancing elements other than H and O
  3. if the reaction takes place in acid or base, balancing O with \( \text{H}_2\text{O} \) and H with \( \text{H}^+ \)
  4. balancing charge with electrons
  5. equalizing the loss and gain of electrons
  6. combining half-reactions, canceling electrons, and combining \( \text{H}_2\text{O} \) and \( \text{H}^+ \)
  7. using \( \text{OH}^- \) to neutralize \( \text{H}^+ \) to \( \text{H}_2\text{O} \) for an oxidation–reduction reaction in basic solution

**15.3 Electrical Energy from Oxidation–Reduction Reactions**

**LEARNING GOAL** Use the activity series to determine if an oxidation–reduction reaction is spontaneous. Write the half-reactions that occur in a voltaic cell and the cell notation.

- In a voltaic cell, the components of the two half-reactions of a spontaneous oxidation–reduction reaction are placed in separate containers called half-cells.
- With a wire connecting the half-cells, an electrical current is generated as electrons move from the anode where oxidation takes place to the cathode where reduction takes place.
15.4 Oxidation–Reduction Reactions That Require Electrical Energy

**LEARNING GOAL** Describe the half-cell reactions and the overall reactions that occur in electrolysis.

- The activity series, which lists metals with the most easily oxidized metal at the top, is used to predict the direction of a spontaneous reaction.

**KEY TERMS**

**activity series** A table of half-reactions with the metals that oxidize most easily at the top, and the metals that do not oxidize easily at the bottom.

**anode** The electrode where oxidation takes place.

**cathode** The electrode where reduction takes place.

**electrolysis** The use of electrical energy to run a nonspontaneous oxidation–reduction reaction in an electrolytic cell.

**electrolytic cell** A cell in which electrical energy is used to make a nonspontaneous oxidation–reduction reaction happen.

**half-reaction method** A method of balancing oxidation–reduction reactions in which the half-reactions are balanced separately and then combined to give the complete reaction.

**oxidation** The loss of electrons by a substance.

**oxidation number** A number equal to zero in an element or the charge of a monatomic ion; in molecular compounds and polyatomic ions, oxidation numbers are assigned using a set of rules.

**oxidation–reduction reaction** A reaction in which electrons are transferred from one reactant to another.

**oxidizing agent** The reactant that gains electrons and is reduced.

**reducing agent** The reactant that loses electrons and is oxidized.

**reduction** The gain of electrons by a substance.

**voltaic cell** A type of cell with two compartments that uses spontaneous oxidation–reduction reactions to produce electrical energy.

**CORE CHEMISTRY SKILLS**

*The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.*

**Assigning Oxidation Numbers (15.1)**

- Oxidation numbers are assigned to atoms or ions in a reaction to determine the substance that is oxidized and the substance that is reduced.
- The rules for assigning oxidation numbers are as follows:
  1. for a molecule the sum is 0; for a polyatomic ion the sum is its charge
  2. for an element is 0
  3. for a monatomic ion is its charge
  4. for Group 1A (1) metals is +1, Group 2A (2) metals is +2
  5. for fluorine is −1; for other Group 7A (17) metals is −1, except when combined with O or F
  6. for oxygen is −2, except in O₂ or peroxides
  7. for H in nonmetal compounds is +1; for H with a metal is −1

**Example:** Assign oxidation numbers to each element in the following:

<table>
<thead>
<tr>
<th>Element</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sn^{4+}</td>
<td>0</td>
</tr>
<tr>
<td>Mn</td>
<td>0</td>
</tr>
<tr>
<td>MnO₂</td>
<td>+4</td>
</tr>
<tr>
<td>MgCO₃</td>
<td>+2</td>
</tr>
</tbody>
</table>

**Answer:**

<table>
<thead>
<tr>
<th>Element</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sn^{4+}</td>
<td>+4</td>
</tr>
<tr>
<td>Mn</td>
<td>0</td>
</tr>
<tr>
<td>MnO₂</td>
<td>+4</td>
</tr>
<tr>
<td>MgCO₃</td>
<td>+2</td>
</tr>
</tbody>
</table>

**Using Oxidation Numbers (15.1)**

- An atom is oxidized when its oxidation number increases from reactant to product.
- An atom is reduced when its oxidation number decreases from reactant to product.

- In an electrolytic cell, electrical energy from an external source is used to make reactions take place that are not spontaneous.
- A method called electrolysis is used to plate chrome on hubcaps, zinc on iron, or gold on stainless steel jewelry.

**Example:** Assign oxidation numbers to each element and identify which is oxidized and which is reduced.

Fe₂O₃(s) + 2Al(s) → Al₂O₃(s) + 2Fe(l)

**Answer:**

Fe(3) → Fe(0) Fe is reduced.

Al(0) → Al(+3) Al is oxidized.

**Identifying Oxidizing and Reducing Agents (15.1)**

- In an oxidation–reduction reaction, the substance that is reduced is the oxidizing agent.
- In an oxidation–reduction reaction, the substance that is oxidized is the reducing agent.

**Example:** Identify the oxidizing agent and the reducing agent in the following:

SnO₂(s) + 2H₂(g) → Sn(s) + 2H₂O(l)

**Answer:**

Oxidation numbers are assigned to identify the substance that is oxidized as the reducing agent and the substance that is reduced as the oxidizing agent.

SnO₂(s) + 2H₂(g) → Sn(s) + 2H₂O(l)

Sn(0) is oxidized; SnO₂ is the oxidizing agent.

H(0) → H(+1)

H is oxidized; H₂ is the reducing agent.
Using Half-Reactions to Balance Redox Equations (15.2)

- The half-reaction method for balancing equations involves separating the oxidation–reduction reaction into two half-reactions.
- balancing the elements in each except H and O; then if the reaction takes place in acid or base balancing O with H₂O and H with H⁺
- balancing charge by adding electrons
- multiplying each half-reaction by factors that equalize the loss and gain of electrons
- combining half-reactions and cancelling electrons, identical ions or molecules, and using OH⁻ in base to neutralize H⁺

Example: Use the half-reaction method to balance the following oxidation–reduction equation in acidic solution:

\[ \text{NO}_3^-(aq) + \text{Sn}^{2+}(aq) \rightarrow \text{NO}(g) + \text{Sn}^{4+}(aq) \]

Answer:

1. \[ \text{NO}_3^-(aq) \rightarrow \text{NO}(g) \]
   \[ \text{Sn}^{2+}(aq) \rightarrow \text{Sn}^{4+}(aq) \]
2. \[ 4\text{H}^{+}(aq) + \text{NO}_3^-(aq) \rightarrow \text{NO}(g) + 2\text{H}_2\text{O}(l) \]
   \[ \text{Sn}^{2+}(aq) \rightarrow \text{Sn}^{4+}(aq) \]
3. \[ 3e^- + 4\text{H}^{+}(aq) + \text{NO}_3^-(aq) \rightarrow \text{NO}(g) + 2\text{H}_2\text{O}(l) \]
   \[ \text{Sn}^{2+}(aq) \rightarrow \text{Sn}^{4+}(aq) + 2e^- \]
4. \[ 2 \times [3e^- + 4\text{H}^{+}(aq) + \text{NO}_3^-(aq) \rightarrow \text{NO}(g) + 2\text{H}_2\text{O}(l)] \]
   \[ 6e^- + 8\text{H}^{+}(aq) + 2\text{NO}_3^-(aq) \rightarrow 2\text{NO}(g) + 4\text{H}_2\text{O}(l) \]
   \[ 3 \times [\text{Sn}^{2+}(aq) \rightarrow \text{Sn}^{4+}(aq) + 2e^-] \]
   \[ 3\text{Sn}^{2+}(aq) \rightarrow 3\text{Sn}^{4+}(aq) + 6e^- \]
5. \[ 8\text{H}^{+}(aq) + 2\text{NO}_3^-(aq) + 3\text{Sn}^{2+}(aq) \rightarrow 2\text{NO}(g) + 3\text{Sn}^{4+}(aq) + 4\text{H}_2\text{O}(l) \]

Understanding the Concepts

The chapter sections to review are shown in parentheses at the end of each question.

15.43 Classify each of the following as oxidation or reduction: (15.1)
   a. Electrons are lost.
   b. Reaction of an oxidizing agent.
   c. \( \text{O}_2(g) \rightarrow \text{OH}^-(aq) \)
   d. \( \text{Br}_2(l) \rightarrow 2\text{Br}^-(aq) \)
   e. \( \text{Sn}^{2+}(aq) \rightarrow \text{Sn}^{4+}(aq) \)

15.44 Classify each of the following as oxidation or reduction: (15.1)
   a. Electrons are gained.
   b. Reaction of a reducing agent.
   c. \( \text{Ni}(s) \rightarrow \text{Ni}^{2+}(aq) \)
   d. \( \text{MnO}_4^-(aq) \rightarrow \text{MnO}_2(s) \)
   e. \( \text{Sn}^{4+}(aq) \rightarrow \text{Sn}^{2+}(aq) \)

15.45 Assign oxidation numbers to the elements in each of the following: (15.1)
   a. \( \text{H}_2\text{O}_2 \)
   b. \( \text{Cl}_2\text{O}_7 \)
   c. \( \text{SO}_3^2- \)
   d. \( \text{N}_2\text{H}_4 \)

15.46 Assign oxidation numbers to the elements in each of the following: (15.1)
   a. \( \text{NaVO}_3 \)
   b. \( \text{NH}_4^+ \)
   c. \( \text{N}_2\text{O} \)
   d. \( \text{MnO}_2 \)

15.47 Which of the following are oxidation–reduction reactions? (15.1)
   a. \( \text{ZnO}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{ZnSO}_4(aq) + \text{H}_2\text{O}(l) \)
   b. \( \text{AgNO}_3(aq) + \text{NaCl}(aq) \rightarrow \text{AgCl}(s) + \text{NaNO}_3(aq) \)
   c. \( \text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g) \)

15.48 Which of the following are oxidation-reduction reactions? (15.1)
   a. \( \text{H}_2\text{CO}_3(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) \)
   b. \( \text{C}_2\text{H}_2\text{O}_6(s) + 6\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 6\text{H}_2\text{O}(l) \)
   c. \( \text{MnO}_2(s) + 4\text{HCl}(aq) \rightarrow \text{MnCl}_2(aq) + 2\text{H}_2\text{O}(l) + \text{Cl}_2(g) \)

15.49 For this reaction, identify each of the following: (15.1)
   \[ 2\text{Cr}_2\text{O}_3(s) + 3\text{Si}(s) \rightarrow 4\text{Cr}(s) + 3\text{SiO}_2(s) \]
   a. the substance reduced
   b. the substance oxidized
c. the oxidizing agent
d. the reducing agent

Identifying Spontaneous Reactions (15.3)

- The activity series in Table 15.3 places the metals that are easily oxidized at the top, and the metals that do not oxidize easily at the bottom.
- A metal will oxidize spontaneously when it is combined with the reverse of the half-reaction of any metal below it on the activity series in Table 15.3.

Example: Write a balanced redox equation for the spontaneous reaction for the following half-reactions on the activity series:

\[ \text{Ca}(s) \rightarrow \text{Ca}^{2+}(aq) + 2e^- \]
\[ \text{Ni}^{2+}(aq) + 2e^- \rightarrow \text{Ni}(s) \]

Answer: The spontaneous reaction is a combination of the oxidation half-reaction higher on the activity series with the reverse of the half-reaction below it.

\[ \text{Ca}(s) \rightarrow \text{Ca}^{2+}(aq) + 2e^- \]
\[ \text{Ni}^{2+}(aq) + 2e^- \rightarrow \text{Ni}(s) \]

Spontaneous reaction

Chromium(III) oxide and silicon undergo an oxidation–reduction reaction.
15.50 For this reaction, identify each of the following: (15.1)
\[ \text{C}_2\text{H}_4(g) + 3\text{O}_2(g) \xrightarrow{\Delta} 2\text{CO}_2(g) + 2\text{H}_2\text{O}(g) \]
a. the substance reduced
b. the substance oxidized
c. the oxidizing agent
d. the reducing agent

15.51 Balance each of the following half-reactions in acidic solution: (15.2)
a. \( \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) \)
b. \( \text{SnO}_2^{2-}(aq) \rightarrow \text{SnO}_2^{4-}(aq) \)
c. \( \text{SO}_3^{2-}(aq) \rightarrow \text{SO}_4^{2-}(aq) \)
d. \( \text{NO}_3^-(aq) \rightarrow \text{NO}(g) \)

15.52 Balance each of the following half-reactions in acidic solution: (15.2)
a. \( \text{I}_2(s) \rightarrow \text{I}^-(aq) \)
b. \( \text{MnO}_4^-(aq) \rightarrow \text{Mn}^{2+}(aq) \)
c. \( \text{Br}_2(l) \rightarrow \text{BrO}_3^-(aq) \)
d. \( \text{ClO}_3^-(aq) \rightarrow \text{ClO}_4^-(aq) \)

15.53 Consider the following voltaic cell: (15.3)

![Voltaic Cell Diagram]

15.54 Consider the following voltaic cell: (15.3)

![Voltaic Cell Diagram]

15.55 Which of the following are oxidation–reduction reactions? (15.1)
a. \( \text{AgNO}_3(aq) + \text{NaCl}(aq) \rightarrow \text{AgCl(s) + NaNO}_3(aq) \)
b. \( 6\text{Li}(s) + \text{N}_2(g) \rightarrow 2\text{Li}_3\text{N}(s) \)
c. \( \text{Ni(s) + Pb(NO}_3)_2(aq) \rightarrow \text{Ni(NO}_3)_2(aq) + \text{Pb}(s) \)
d. \( 2\text{K}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{KOH}(aq) + \text{H}_2(g) \)

15.56 Which of the following are oxidation–reduction reactions? (15.1)
a. \( \text{Ca}(s) + \text{F}_2(g) \rightarrow \text{CaF}_2(s) \)
b. \( \text{Fe}(s) + 2\text{HCl}(aq) \rightarrow \text{FeCl}_2(aq) + \text{H}_2(g) \)
c. \( 2\text{NaCl}(aq) + \text{Pb(NO}_3)_2(aq) \rightarrow \text{PbCl}_2(s) + 2\text{NaNO}_3(aq) \)
d. \( 2\text{CuCl}(aq) \rightarrow \text{Cu}(s) + \text{CuCl}_2(aq) \)

15.57 Pure iron can be produced from iron oxide in a blast furnace by the following reaction: (15.1)
\[ 3\text{C}(s) + 2\text{Fe}_2\text{O}_3(s) \rightarrow 4\text{Fe}(s) + 3\text{CO}_2(g) \]
a. Which is the oxidizing agent?
b. Which is the reducing agent?

15.58 Household bleach is a solution containing 3–8% sodium hypochlorite (\( \text{NaOCl} \)) to kill microbes. If the product is \( \text{Cl}^- \), was the elemental chlorine oxidized or reduced? (15.1)

15.59 Assign oxidation numbers to all the elements in each of the following: (15.1)
a. \( \text{K}_2\text{Cr}_2\text{O}_7 \)
b. \( \text{AsO}_3^{3-} \)
c. \( \text{F}_2\text{O} \)
d. \( \text{FeSO}_3 \)
e. \( \text{CS}_2 \)

15.60 Assign oxidation numbers to all the elements in each of the following: (15.1)
a. \( \text{WF}_6 \)
b. \( \text{VN} \)
c. \( \text{PtCl}_4^{2-} \)
d. \( \text{Ag}_2\text{S} \)
e. \( \text{Cu(NO}_3)_2 \)

15.61 Assign oxidation numbers to all the elements in each of the following reactions, and identify the reactant that is oxidized, the reactant that is reduced, the oxidizing agent, and the reducing agent: (15.1)
a. \( 2\text{FeCl}_2(aq) + \text{Cl}_2(g) \rightarrow 2\text{FeCl}_3(aq) \)
b. \( 2\text{H}_2\text{S}(g) + 3\text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) + 2\text{SO}_2(g) \)
c. \( \text{P}_2\text{O}_5(s) + 3\text{C}(s) \rightarrow 2\text{P}(s) + 5\text{CO}(g) \)

15.62 Assign oxidation numbers to all the elements in each of the following reactions, and identify the reactant that is oxidized, the reactant that is reduced, the oxidizing agent, and the reducing agent: (15.1)
a. \( 4\text{Al}(s) + 3\text{O}_2(g) \rightarrow 2\text{Al}_2\text{O}_3(s) \)
b. \( \text{I}_2\text{O}_5(s) + 5\text{CO}(g) \rightarrow \text{I}_2(g) + 5\text{CO}_2(g) \)
c. \( 2\text{Cr}_2\text{O}_3(s) + 3\text{C}(s) \rightarrow 4\text{Cr}(s) + 3\text{CO}_2(g) \)

15.63 Write the balanced half-reactions and a balanced redox equation for each of the following reactions in acidic solution: (15.2)
a. \( \text{Zn}(s) + \text{NO}_3^-(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{NO}_2(g) \)
b. \( \text{MnO}_4^-(aq) + \text{SO}_3^{2-}(aq) \rightarrow \text{Mn}^{2+}(aq) + \text{SO}_4^{2-}(aq) \)
c. \( \text{ClO}_3^-(aq) + \text{I}^-(aq) \rightarrow \text{Cl}^-(aq) + \text{I}_2(s) \)
d. \( \text{Cr}_2\text{O}_7^{2-}(aq) + \text{C}_2\text{O}_4^{2-}(aq) \rightarrow \text{Cr}^{3+}(aq) + \text{CO}_2(g) \)
15.64 Write the balanced half-reactions and a balanced redox equation for each of the following reactions in acidic solution: (15.2)
   a. Sn^{2+}(aq) + 2H^+(aq) \rightarrow Sn^{4+}(aq) + H_2(g)
   b. S_2O_3^{2-}(aq) + H_2O(l) \rightarrow S_2O_4^{2-}(aq) + 2H^+(aq)
   c. Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)
   d. Al(s) + Cr_2O_7^{2-}(aq) \rightarrow Al^{3+}(aq) + Cr^{3+}(aq)

15.65 Use the activity series in Table 15.3 to predict whether each of the following reactions will occur spontaneously: (15.3)
   a. 2Cr(s) + 3Ni^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Ni(s)
   b. Cu(s) + Zn^{2+}(aq) \rightarrow Cu^{2+}(aq) + Zn(s)
   c. Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)

15.66 Use the activity series in Table 15.3 to predict whether each of the following reactions will occur spontaneously: (15.3)
   a. Zn(s) + Mg^{2+}(aq) \rightarrow Zn^{2+}(aq) + Mg(s)
   b. 3Na(s) + Al^{3+}(aq) \rightarrow 3Na^+(aq) + Al(s)
   c. Mg(s) + Ni^{2+}(aq) \rightarrow Mg^{2+}(aq) + Ni(s)

15.67 For the voltaic cell Mg(s) \| Mg^{2+}(aq) \| Al^{3+}(aq) \| Al(s), identify each of the following: (15.3)
   a. the anode
   b. the cathode
   c. the half-reaction at the anode
   d. the half-reaction at the cathode
   e. the overall reaction
   f. a solution that can be used for the salt bridge

15.68 For the voltaic cell Al(s) \| Al^{3+}(aq) \| Sn^{2+}(aq) \| S(s), identify each of the following: (15.3)
   a. the anode
   b. the cathode
   c. the half-reaction at the anode
   d. the half-reaction at the cathode
   e. the overall reaction
   f. a solution that can be used for the salt bridge

15.69 Use the activity series in Table 15.3 to determine which of the following ions will be reduced when an iron strip is placed in an aqueous solution of that ion: (15.3)
   a. Cu^{2+}(aq)
   b. Ag^{+}(aq)
   c. Ni^{2+}(aq)
   d. Al^{3+}(aq)
   e. Pb^{2+}(aq)

15.70 Use the activity series in Table 15.3 to determine which of the following ions will be reduced when an aluminum strip is placed in an aqueous solution of that ion: (15.3)
   a. Fe^{2+}(aq)
   b. Au^{3+}(aq)
   c. Mg^{2+}(aq)
   d. H^{+}(aq)
   e. Pb^{2+}(aq)

15.71 In a lead storage battery, the following unbalanced half-reaction takes place: (15.3)
   Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)
   a. Balance the half-reaction.
   b. Is Pb(s) oxidized or reduced?
   c. Indicate whether the half-reaction takes place at the anode or cathode.

15.72 In an acidic dry-cell battery, the following unbalanced half-reaction takes place in acidic solution: (15.3)
   MnO_2(s) \rightarrow Mn_2O_3(s)
   a. Balance the half-reaction.
   b. Is MnO_2(s) oxidized or reduced?
   c. Indicate whether the half-reaction takes place at the anode or cathode.

15.73 Steel bolts made for sailboats are coated with zinc. Add the necessary components (electrodes, wires, batteries) to this diagram of an electrolytic cell with a zinc nitrate solution to show how it could be used to zinc plate a steel bolt. (15.3, 15.4)

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\[ Zn^{2+}(aq) \rightarrow Zn(s) \]\n\[ 2NO_3^{-}(aq) \rightarrow NO_2(g) + O_2(g) + 4H^+(aq) + 2e^- \]
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   a. What is the anode?
   b. What is the cathode?
   c. What is the half-reaction that takes place at the anode?
   d. What is the half-reaction that takes place at the cathode?
   e. If steel is mostly iron, what is the purpose of the zinc coating?

15.74 Copper cooking pans are stainless steel pans plated with a layer of copper. Add the necessary components (electrodes, wires, batteries) to this diagram of an electrolytic cell with a copper(II) nitrate solution to show how it could be used to copper plate a stainless steel (iron) pan. (15.3, 15.4)

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\[ Cu^{2+}(aq) \rightarrow Cu(s) \]\n\[ 2NO_3^{-}(aq) \rightarrow NO_2(g) + O_2(g) + 4H^+(aq) + 2e^- \]
```

   a. What is the anode?
   b. What is the cathode?
   c. What is the half-reaction that takes place at the anode?
   d. What is the half-reaction that takes place at the cathode?
**CHALLENGE QUESTIONS**

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

15.75 Determine the oxidation number of I in each of the following: (15.1)
   a. I2
   b. KI
   c. HIO3
   d. H3IO6

15.76 Determine the oxidation number of P in each of the following: (15.1)
   a. P4
   b. PCl3
   c. PF6−
   d. H3PO4

15.77 Use half-reactions to balance the following equation in acidic solution: (15.2)
   \[
   \text{Cr}_2\text{O}_7^{2−}(aq) + \text{NO}_2^{−}(aq) \rightarrow \text{Cr}^{3+}(aq) + \text{NO}_3^{−}(aq)
   \]

15.78 Use half-reactions to balance the following equation in acidic solution: (15.2)
   \[
   \text{Mn}(s) + \text{Cr}^{3+}(aq) \rightarrow \text{Mn}^{2+}(aq) + \text{Cr}(s)
   \]

15.79 The following unbalanced reaction takes place in acidic solution: (11.6, 15.2)
   \[
   \text{Ag}(s) + \text{NO}_3^{−}(aq) \rightarrow \text{Ag}^+(aq) + \text{NO}(g)
   \]
   a. Write the balanced equation.
   b. How many liters of NO(g) are produced at STP when 15.0 g of silver react with excess nitric acid?

15.80 The following unbalanced reaction takes place in acidic solution: (12.6, 15.2)
   \[
   \text{MnO}_4^{−}(aq) + \text{Fe}^{2+}(aq) \rightarrow \text{Mn}^{2+}(aq) + \text{Fe}^{3+}(aq)
   \]
   a. Write the balanced equation.
   b. How many milliliters of a 0.150 M KMnO4 solution are needed to react with 25.0 mL of a 0.400 M FeSO4 solution?

15.81 A concentrated nitric acid solution is used to dissolve copper(II) sulfide. (12.6, 15.2)
   \[
   \text{CuS}(s) + \text{HNO}_3(aq) \rightarrow \text{CuSO}_4(aq) + \text{NO}(g) + \text{H}_2\text{O}(l)
   \]
   a. Write the balanced equation.
   b. How many milliliters of a 16.0 M HNO3 solution are needed to dissolve 24.8 g of CuS?

15.82 The following unbalanced reaction takes place in acidic solution: (12.6, 15.2)
   \[
   \text{Cr}_2\text{O}_7^{2−}(aq) + \text{Fe}^{2+}(aq) \rightarrow \text{Cr}^{3+}(aq) + \text{Fe}^{3+}(aq)
   \]
   a. Write the balanced equation.
   b. How many milliliters of a 0.211 M K2Cr2O7 solution are needed to react with 5.00 g of FeSO4?

15.83 Balance the equation for the reaction that takes place in basic solution. (15.2)
   \[
   \text{Se}(s) + \text{Cr(OH)}_3(aq) \rightarrow \text{Cr}(s) + \text{SeO}_3^{2−}(aq)
   \]

15.84 Balance the equation for the reaction that takes place in basic solution. (15.2)
   \[
   \Gamma(aq) + \text{OCl}^{−}(aq) \rightarrow \text{I}_2(s) + \text{Cl}^{−}(aq)
   \]

15.85 Using the activity series in Table 15.3, indicate whether each of the following reactions is spontaneous: (15.3)
   a. Zn(s) + Cu2+(aq) \rightarrow Zn2+(aq) + Cu(s)
   b. 2Al(s) + 3Sn2+(aq) \rightarrow 2Al3+(aq) + 3Sn(s)
   c. Mg(s) + 2H+(aq) \rightarrow Mg2+(aq) + H2(g)

15.86 Using the activity series in Table 15.3, indicate whether each of the following reactions is spontaneous: (15.3)
   a. Cu(s) + Ni2+(aq) \rightarrow Cu2+(aq) + Ni(s)
   b. 2Cr(s) + 3Fe2+(aq) \rightarrow 2Cr3+(aq) + 3Fe(s)
   c. Fe(s) + Mg2+(aq) \rightarrow Fe2+(aq) + Mg(s)

15.87 Draw a diagram of a voltaic cell for Ni(s) ∥ Ag+(aq) ∥ Ag(s). (15.3)
   a. What is the anode?
   b. What is the cathode?
   c. What is the half-reaction that takes place at the anode?
   d. What is the half-reaction that takes place at the cathode?
   e. What is the overall reaction for the cell?

15.88 Draw a diagram of a voltaic cell for Mg(s) ∥ Mg2+(aq) ∥ Al3+(aq) ∥ Al(s). (15.3)
   a. What is the anode?
   b. What is the cathode?
   c. What is the half-reaction that takes place at the anode?
   d. What is the half-reaction that takes place at the cathode?
   e. What is the overall reaction for the cell?

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**ANSWERS**

Answers to Selected Questions and Problems

15.1 a. reduction   b. oxidation   c. reduction

15.3 a. Al is oxidized and O₂ is reduced.
   b. Zn is oxidized and H⁺ is reduced.
   c. Br⁻ is oxidized and F₂ is reduced.

15.5 a. 0   b. 0   c. +2   d. −1

15.7 a. K is +1, Cl is −1.
   b. Mn is +4, O is −2.
   c. N is +2, O is −2.
   d. Mn is +3, O is −2.

15.9 a. Li is +1, P is +5, and O is −2.
   b. S is +4, O is −2.
   c. Cr is +3, S is −2.
   d. N is +5, O is −2.

15.11 a. H is +1, S is +6, and O is −2.
   b. H is +1, P is +3, and O is −2.
   c. Cr is +6, O is −2.
   d. Na is +1, C is +4, and O is −2.

15.13 a. +5   b. −2   c. +5   d. +6

15.15 a. The substance that is oxidized is the reducing agent.
   b. The substance that gains electrons is reduced and is the oxidizing agent.

15.17 a. Li is oxidized; Li is the reducing agent. Cl (in Cl₂) is reduced; Cl₂ is the oxidizing agent.
   b. Br⁻ (in NaBr) is oxidized; NaBr is the reducing agent. Cl (in Cl₂) is reduced; Cl₂ is the oxidizing agent.
   c. Pb is oxidized; Pb is the reducing agent. O (in O₂) is reduced; O₂ is the oxidizing agent.
   d. Al is oxidized; Al is the reducing agent. Ag⁺ is reduced; Ag⁻ is the oxidizing agent.
15.19 a. $S^{2-}$ (in NiS) is oxidized; NiS is the reducing agent.
O (in $O_2$) is reduced; $O_2$ is the oxidizing agent.
b. Sn$^{2+}$ is oxidized; Sn$^{2+}$ is the reducing agent. Fe$^{3+}$ is reduced; Fe$^{3+}$ is the oxidizing agent.
c. C (in CH$_4$) is oxidized; CH$_4$ is the reducing agent. O (in $O_2$) is reduced; O$_2$ is the oxidizing agent.
d. Si is oxidized; Si is the reducing agent. Cr$^{3+}$ (in Cr$_2$O$_3$) is reduced; Cr$_2$O$_3$ is the oxidizing agent.

15.21 a. Sn$^{2+}$(aq) $\rightarrow$ Sn$^{4+}$(aq) + 2 $e^-$
b. Mn$^{2+}$(aq) + 4H$_2$O(l) $\rightarrow$ MnO$_4^-$ (aq) + 8H$^+$ (aq) + 5 $e^-$
c. NO$_3^-$ (aq) + H$_2$O(l) $\rightarrow$ NO$_2^-$ (aq) + 2H$^+$ (aq) + 2 $e^-$
d. $e^-$ + 2H$^+$ (aq) + ClO$_3^-$ (aq) $\rightarrow$ ClO$_2$(aq) + H$_2$O(l)

15.23 a. 2H$^+$ (aq) + Ag(s) + NO$_3^-$ (aq) $\rightarrow$ Ag$^+$ (aq) + NO$_2$(g) + H$_2$O(l)
   b. 4H$^+$ (aq) + 4NO$^-$(aq) + 3S(s) $\rightarrow$ 4NO(g) + 3SO$_2$(g) + 2H$_2$O(l)
   c. 2S$_2$O$_3^{2-}$(aq) + Cu$^{2+}$(aq) $\rightarrow$ S$_4$O$_6^{2-}$(aq) + Cu(s)

15.25 a. 2Fe(s) + 2CrO$_4^{2-}$(aq) + 2H$_2$O(l) $\rightarrow$ Fe$_2$O$_3(s)$ + Cr$_2$O$_3(s)$ + 4OH$^-$ (aq)
   b. 3CN$^-$ (aq) + 2MnO$_4^-$ (aq) + H$_2$O(l) $\rightarrow$ 3CNO$^-$(aq) + 2MnO$_2$(s) + 2OH$^-$ (aq)

15.27 a. Since Au is below H$_2$ in the activity series, the reaction will not be spontaneous.
b. Since Fe is above Ni in the activity series, the reaction will not be spontaneous.
c. Since Ag is below Cu in the activity series, the reaction will not be spontaneous.

15.29 a. Anode reaction: Pb(s) $\rightarrow$ Pb$^{2+}$(aq) + 2 $e^-$
   Cathode reaction: Cu$^{2+}$(aq) + 2 $e^-$ $\rightarrow$ Cu(s)
   Overall cell reaction: Pb(s) + Cu$^{2+}$(aq) $\rightarrow$ Pb$^{2+}$(aq) + Cu(s)
b. Anode reaction: Cr(s) $\rightarrow$ Cr$^{3+}$(aq) + 2 $e^-$
   Cathode reaction: Ag$^+$ (aq) + $e^-$ $\rightarrow$ Ag(s)
   Overall cell reaction: Cr(s) + 2Ag$^+$ (aq) $\rightarrow$ Cr$^{3+}$(aq) + 2Ag(s)

15.31 a. The anode is a Cd metal electrode in a Cd$^{2+}$ solution. The anode reaction is
   Cd(s) $\rightarrow$ Cd$^{2+}$(aq) + 2 $e^-$
   The cathode is a Sn metal electrode in a Sn$^{2+}$ solution. The cathode reaction is
   Sn$^{2+}$(aq) + 2 $e^-$ $\rightarrow$ Sn(s)
   The shorthand notation for this cell is
   Cd(s) $\rightarrow$ Cd$^{2+}$(aq) + Sn(s)
b. The anode is a Zn metal electrode in a Zn$^{2+}$ solution. The anode reaction is
   Zn(s) $\rightarrow$ Zn$^{2+}$(aq) + 2 $e^-$
   The cathode is a C (graphite) electrode, where Cl$_2$ gas is reduced to Cl$^-$. The cathode reaction is
   Cl$_2$(g) + 2 $e^-$ $\rightarrow$ 2Cl$^-$ (aq)
   The shorthand notation for this cell is
   Zn(s) $\rightarrow$ Zn$^{2+}$(aq) $\parallel$ Cl$_2$(g), Cl$^-$ (aq) $\parallel$ C

15.33 a. The half-reaction is an oxidation.
b. Cd metal is oxidized.
c. Oxidation takes place at the anode.

15.35 a. The half-reaction is an oxidation.
b. Zn metal is oxidized.
c. Oxidation takes place at the anode.

15.37 a. Sn$^{2+}$(aq) + 2 $e^-$ $\rightarrow$ Sn(s)
b. The reduction of Sn$^{2+}$ to Sn occurs at the cathode, which is the iron can.
c. The oxidation of Sn to Sn$^{2+}$ occurs at the anode, which is the tin bar.

15.39 Since Fe is above Sn in the activity series, if the Fe is exposed to air and water, Fe will be oxidized and rust will form. To protect iron, Sn would have to be more active than Fe and it is not.

15.41 a. H$_2$ is oxidized and O$_2$ is reduced.
b. O$_2$ is the oxidizing agent.
c. H$_2$ is the reducing agent.
d. H$_2$(g) + O$_2$(g) $\rightarrow$ H$_2$O$_2$(aq)

15.43 a. oxidation b. reduction
   c. reduction d. reduction
e. oxidation

15.45 a. H = +1, O = −1
   b. O = −2, Cl = +7
   c. O = −2, S = +4
   d. H = +1, N = −2

15.47 Reaction c involves loss and gain of electrons; c is an oxidation–reduction reaction.

15.49 a. Cr in Cr$_2$O$_3$ is reduced.
b. Si is oxidized.
c. Cr$_2$O$_3$ is the oxidizing agent.
d. Si is the reducing agent.

15.51 a. Zn(s) $\rightarrow$ Zn$^{2+}$(aq) + 2 $e^-$
b. SnO$_2^{2-}$(aq) + H$_2$O(l) $\rightarrow$ SnO$_3^{2-}$(aq) + 2H$^+$ (aq) + 2 $e^-$
c. SO$_3^{2-}$(aq) + H$_2$O(l) $\rightarrow$ SO$_4^{2-}$(aq) + 2H$^+$ (aq) + 2 $e^-$
d. 3 $e^-$ + 4H$^+$ (aq) + NO$_3^-$ (aq) $\rightarrow$ NO(g) + 2H$_2$O(l)

15.53 a. Fe(s) $\rightarrow$ Fe$^{2+}$(aq) + 2 $e^-$
b. Ni$^{2+}$(aq) + 2 $e^-$ $\rightarrow$ Ni(s)
c. Fe is the anode.
d. Ni is the cathode.

15.55 Reactions b, c, and d all involve loss and gain of electrons; b, c, and d are oxidation–reduction reactions.

15.57 a. Fe is gaining electrons; this is a reduction.
b. C is losing electrons; this is an oxidation.

15.59 a. O = −2, K = +1, Cr = +6
   b. O = −2, As = +5
   c. F = −1, O = +2
   d. Fe = +2, S = +6, O = −2
   e. S = −2, C = +4

15.61 a. 2FeCl$_3$(aq) + Cl$_2$(g) $\rightarrow$ 2FeCl$_2$(aq)
   Fe in FeCl$_3$ is oxidized; FeCl$_3$ is the reducing agent.
   Cl in Cl$_2$ is reduced; Cl$_2$ is the oxidizing agent.
   b. 2H$_2$S(g) + $3O_2$(g) $\rightarrow$ 2H$_2$O(l) + $2SO_2$(g)
   S in H$_2$S is oxidized; H$_2$S is the reducing agent.
   O in $O_2$ is reduced; $O_2$ is the oxidizing agent.
   c. P$_2$O$_5$(s) + 5C(s) $\rightarrow$ 2P(s) + 5CO(g)
   C is oxidized; C is the reducing agent.
   P in P$_2$O$_5$ is reduced; P$_2$O$_5$ is the oxidizing agent.
15.63 a. Zn(s) → Zn^{2+}(aq) + 2e^{-};
   Overall:
   \[ 4\text{H}^+(aq) + \text{Zn(s)} + 2\text{NO}_3^-(aq) \rightarrow \text{Zn}^{2+}(aq) + 2\text{NO}_2(g) + 2\text{H}_2\text{O(l)} \]
   b. 5e^{-} + 8\text{H}^+(aq) + \text{MnO}_4^-(aq) → Mn^{2+}(aq) + 4\text{H}_2\text{O(l)};
   Overall:
   \[ 6\text{H}^+(aq) + 2\text{MnO}_4^-(aq) + 5\text{SO}_3^{2-}(aq) \rightarrow 2\text{Mn}^{2+}(aq) + 5\text{SO}_4^{2-}(aq) + 3\text{H}_2\text{O(l)} \]

15.65 a. Since Cr is above Ni in the activity series, the reaction will be spontaneous.
   b. Since Cu is below Zn in the activity series, the reaction will not be spontaneous.
   c. Since Zn is above Pb in the activity series, the reaction will be spontaneous.

15.67 a. The anode is Mg.
   b. The cathode is Al.
   c. The half-reaction at the anode is
   \[ \text{Mg}(s) \rightarrow \text{Mg}^{2+}(aq) + 2\text{e}^{-} \]
   d. The half-reaction at the cathode is
   \[ \text{Al}^{3+}(aq) + 3\text{e}^{-} \rightarrow \text{Al(s)} \]
   e. The overall reaction is
   \[ 3\text{Mg}(s) + 2\text{Al}^{3+}(aq) \rightarrow 3\text{Mg}^{2+}(aq) + 2\text{Al(s)} \]
   f. An NaCl solution can be used for the salt bridge.

15.69 a. Ca^{2+}(aq) will not be reduced by an iron strip.
   b. Ag^{+}(aq) will be reduced by an iron strip.
   c. Ni^{2+}(aq) will be reduced by an iron strip.
   d. Al^{3+}(aq) will not be reduced by an iron strip.
   e. Pb^{2+}(aq) will be reduced by an iron strip.

15.71 a. Pb(s) + SO_3^{2-}(aq) → PbSO_4(s) + 2e^{-}
   b. Pb(s) is oxidized.
   c. The half-reaction takes place at the anode.

15.73 a. The anode is a bar of zinc.
   b. The cathode is the steel bolt.
   c. The half-reaction at the anode is
   \[ \text{Zn(s)} \rightarrow \text{Zn}^{2+}(aq) + 2\text{e}^{-} \]
   d. The half-reaction at the cathode is
   \[ \text{Zn}^{2+}(aq) + 2\text{e}^{-} \rightarrow \text{Zn(s)} \]
   e. The purpose of the zinc coating is to prevent rusting of the bolt by H_2O and O_2.

15.75 a. 0
   b. -1
   c. +5
   d. +7

15.77 a. \[ 8\text{H}^+(aq) + \text{Cr}_2\text{O}_7^{2-}(aq) + 3\text{NO}_2^-(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 3\text{NO}_3^-(aq) + 4\text{H}_2\text{O(l)} \]

15.79 a. \[ 4\text{H}^+(aq) + 3\text{Ag(s)} + \text{NO}_3^-(aq) \rightarrow 3\text{Ag}^{+}(aq) + \text{NO(g)} + 2\text{H}_2\text{O(l)} \]

15.81 a. \[ 3\text{CuS}(s) + 8\text{HNO}_3(aq) \rightarrow 3\text{CuSO}_4(aq) + 8\text{NO(g)} + 4\text{H}_2\text{O(l)} \]
   b. 43.2 mL of HNO_3 solution

15.83 \[ 3\text{Se}(s) + 4\text{Cr(OH)}_3(aq) + 6\text{H}^+(aq) \rightarrow 4\text{Cr}(s) + 3\text{SeO}_3^2-(aq) + 9\text{H}_2\text{O(l)} \]

15.85 a. Since Zn is below Ca in the activity series, the reaction will not be spontaneous.
   b. Since Al is above Sn in the activity series, the reaction will be spontaneous.
   c. Since Mg is above H_2 in the activity series, the reaction will be spontaneous.

15.87

\[ \text{Ni(s)} \]
\[ \text{Salt bridge} \]
\[ \text{Ag(s)} \]
SIMONE’S DOCTOR IS concerned about her elevated cholesterol, which could lead to coronary heart disease and a heart attack. He sends her to a nuclear medicine center to undergo a cardiac stress test. Dr. Paul, the radiologist, explains to Simone that a stress test measures the blood flow to her heart muscle at rest and then during stress. The test is performed in a similar way as a routine exercise stress test, but images are produced that show areas of low blood flow through the heart and areas of damaged heart muscle. It involves taking two sets of images of her heart, one when the heart is at rest, and one when she is exercising on a treadmill.

Dr. Paul tells Simone that he will inject thallium-201 into her bloodstream. He explains that TI-201 is a radioactive isotope that has a half-life of 3.0 days. Simone is curious about the term “half-life.”

Dr. Paul explains that a half-life is the amount of time it takes for one-half of a radioactive sample to break down. He assures her that after four half-lives, the radiation emitted will be almost zero. Dr. Paul tells Simone that TI-201 decays to Hg-201 and emits energy similar to X-rays. When the TI-201 reaches any areas within her heart with restricted blood supply, smaller amounts of the radioisotope will accumulate.

CAREER

Radiologist

A radiologist works in a hospital or imaging center where nuclear medicine is used to diagnose and treat a variety of medical conditions. In a diagnostic test, the radiologist uses a scanner, which converts radiation into images. The images are evaluated to determine any abnormalities in the body. A radiologist operates the instrumentation and computers associated with nuclear medicine such as computed tomography (CT), magnetic resonance imaging (MRI), and positron emission tomography (PET). A radiologist must know how to handle radioisotopes safely, use the necessary type of shielding, and give radioactive isotopes to patients. In addition, they must physically and mentally prepare patients for imaging. A patient may be given radioactive tracers such as technetium-99m, iodine-131, gallium-67, and thallium-201 that emit gamma radiation, which is detected and used to develop an image of the kidneys or thyroid or to follow the blood flow in the heart muscle.
16.1 **Natural Radioactivity**

**LEARNING GOAL** Describe alpha, beta, positron, and gamma radiation.

Most naturally occurring isotopes of elements up to atomic number 19 have stable nuclei. Elements with atomic numbers 20 and higher usually have one or more isotopes that have unstable nuclei in which the nuclear forces cannot offset the repulsions between the protons. An unstable nucleus is radioactive, which means that it spontaneously emits small particles of energy called radiation to become more stable. Radiation may take the form of alpha (α) and beta (β) particles, positrons (β⁺), or pure energy such as gamma (γ) rays. An isotope of an element that emits radiation is called a radioisotope. For most types of radiation, there is a change in the number of protons in the nucleus, which means that an atom is converted into an atom of a different element. This kind of nuclear change was not evident to Dalton when he made his predictions about atoms. Elements with atomic numbers of 93 and higher are produced artificially in nuclear laboratories and consist only of radioactive isotopes.

The atomic symbols for the different isotopes are written with the mass number in the upper left corner and the atomic number in the lower left corner. The mass number is the sum of the number of protons and neutrons in the nucleus, and the atomic number is equal to the number of protons. For example, a radioactive isotope of iodine used in the diagnosis and treatment of thyroid conditions has a mass number of 131 and an atomic number of 53.

![Symbol of element](image)

Mass number (protons and neutrons)  
Atomic number (protons)

Radioactive isotopes are identified by writing the mass number after the element’s name or symbol. Thus, in this example, the isotope is called iodine-131 or I-131.

**TABLE 16.1** compares some stable, nonradioactive isotopes with some radioactive isotopes.

<table>
<thead>
<tr>
<th>Magnesium</th>
<th>Iodine</th>
<th>Uranium</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Stable Isotopes</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>²⁴Mg</td>
<td>¹²⁷I</td>
<td>None</td>
</tr>
<tr>
<td>Magnesium-24</td>
<td>Iodine-127</td>
<td></td>
</tr>
<tr>
<td><strong>Radioactive Isotopes</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>²⁴Mg</td>
<td>¹²⁵I</td>
<td>²³⁵U</td>
</tr>
<tr>
<td>Magnesium-23</td>
<td>Iodine-125</td>
<td>Uranium-235</td>
</tr>
<tr>
<td>²⁷Mg</td>
<td>¹³¹I</td>
<td>²³⁸U</td>
</tr>
<tr>
<td>Magnesium-27</td>
<td>Iodine-131</td>
<td>Uranium-238</td>
</tr>
</tbody>
</table>
Types of Radiation

By emitting radiation, an unstable nucleus forms a more stable, lower energy nucleus. One type of radiation consists of alpha particles. An alpha particle is identical to a helium (He) nucleus, which has two protons and two neutrons. An alpha particle has a mass number of 4, an atomic number of 2, and a charge of 2+. The symbol for an alpha particle is the Greek letter alpha (α) or the symbol of a helium nucleus except that the 2+ charge is omitted.

Another type of radiation occurs when a radioisotope emits a beta particle. A beta particle, which is a high-energy electron, has a charge of 1−, and because its mass is so much less than the mass of a proton, it has a mass number of 0. It is represented by the Greek letter beta (β) or by the symbol for the electron including the mass number and the charge (\( _0^1e \)). A beta particle is formed when a neutron in an unstable nucleus changes into a proton.

A positron, similar to a beta particle, has a positive (1+) charge with a mass number of 0. It is represented by the Greek letter beta with a 1+ charge, β⁰⁺, or by the symbol for an electron, which includes the mass number and the charge (\( _0^1e \)). A positron is produced by an unstable nucleus when a proton is transformed into a neutron and a positron.

A positron is an example of antimatter, a term physicists use to describe a particle that is the opposite of another particle, in this case, an electron. When an electron and a positron collide, their minute masses are completely converted to energy in the form of gamma rays:

\[ -_0^1e + _0^1e \rightarrow 2^0\gamma \]

Gamma rays are high-energy radiation, released when an unstable nucleus undergoes a rearrangement of its particles to give a more stable, lower energy nucleus. Gamma rays are often emitted along with other types of radiation. A gamma ray is written as the Greek letter gamma (\( ^0\gamma \)). Because gamma rays are energy only, zeros are used to show that a gamma ray has no mass or charge.

TABLE 16.2 summarizes the types of radiation we will use in nuclear equations.

<table>
<thead>
<tr>
<th>Type of Radiation</th>
<th>Symbol</th>
<th>Change in Nucleus</th>
<th>Mass Number</th>
<th>Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alpha Particle</td>
<td>α</td>
<td>( \frac{4}{2}\text{He} )</td>
<td>Two protons and two neutrons are emitted as an alpha particle.</td>
<td>4</td>
</tr>
<tr>
<td>Beta Particle</td>
<td>β</td>
<td>( _0^1e )</td>
<td>A neutron changes to a proton and an electron is emitted.</td>
<td>0</td>
</tr>
<tr>
<td>Positron</td>
<td>β⁺</td>
<td>( _0^1e )</td>
<td>A proton changes to a neutron and a positron is emitted.</td>
<td>0</td>
</tr>
<tr>
<td>Gamma Ray</td>
<td>γ</td>
<td>( _0^0\gamma )</td>
<td>Energy is lost to stabilize the nucleus.</td>
<td>0</td>
</tr>
<tr>
<td>Proton</td>
<td>p</td>
<td>( _1^1\text{H} )</td>
<td>A proton is emitted.</td>
<td>1</td>
</tr>
<tr>
<td>Neutron</td>
<td>n</td>
<td>( _1^0\text{n} )</td>
<td>A neutron is emitted.</td>
<td>1</td>
</tr>
</tbody>
</table>

SAMPLE PROBLEM 16.1 Radiation Particles

Identify and write the symbol for each of the following types of radiation:

a. contains two protons and two neutrons
b. has a mass number of 0 and a 1− charge

TRY IT FIRST

SOLUTION

a. An alpha (α) particle, \( \frac{4}{2}\text{He} \), has two protons and two neutrons.
b. A beta (β) particle, \( _0^1e \), is like an electron with a mass number of 0 and a 1− charge.
Biological Effects of Radiation

When radiation strikes molecules in its path, electrons may be knocked away, forming unstable ions. If this ionizing radiation passes through the human body, it may interact with water molecules, removing electrons and producing H_2O^+, which can cause undesirable chemical reactions.

The cells most sensitive to radiation are the ones undergoing rapid division—those of the bone marrow, skin, reproductive organs, and intestinal lining, as well as all cells of growing children. Damaged cells may lose their ability to produce necessary materials. For example, if radiation damages cells of the bone marrow, red blood cells may no longer be produced. If sperm cells, ova, or the cells of a fetus are damaged, birth defects may result. In contrast, cells of the nerves, muscles, liver, and adult bones are much less sensitive to radiation because they undergo little or no cellular division.

Cancer cells are another example of rapidly dividing cells. Because cancer cells are highly sensitive to radiation, large doses of radiation are used to destroy them. The normal tissue that surrounds cancer cells divides at a slower rate and suffers less damage from radiation. However, radiation may cause malignant tumors, leukemia, anemia, and genetic mutations.

Radiation Protection

Nuclear medicine technologists, chemists, doctors, and nurses who work with radioactive isotopes must use proper radiation protection. Proper shielding is necessary to prevent exposure. Alpha particles, which have the largest mass and charge of the radiation particles, travel only a few centimeters in the air before they collide with air molecules, acquire electrons, and become helium atoms. A piece of paper, clothing, and our skin are protection against alpha particles. Lab coats and gloves will also provide sufficient shielding. However, if alpha emitters are ingested or inhaled, the alpha particles they give off can cause serious internal damage.

Beta particles have a very small mass and move much faster and farther than alpha particles, traveling as much as several meters through air. They can pass through paper and penetrate as far as 4 to 5 mm into body tissue. External exposure to beta particles can burn the surface of the skin, but they do not travel far enough to reach the internal organs. Heavy clothing such as lab coats and gloves are needed to protect the skin from beta particles.

Gamma rays travel great distances through the air and pass through many materials, including body tissues. Because gamma rays penetrate so deeply, exposure to gamma rays can be extremely hazardous. Only very dense shielding, such as lead or concrete, will stop them. Syringes used for injections of radioactive materials use shielding made of lead or heavy-weight materials such as tungsten and plastic composites.

When working with radioactive materials, medical personnel wear protective clothing and gloves and stand behind a shield (see Figure 16.1). Long tongs may be used to pick up vials of radioactive material, keeping them away from the hands and body.

Table 16.3 summarizes the shielding materials required for the various types of radiation.
If you work in an environment where radioactive materials are present, such as a nuclear medicine facility, try to keep the time you spend in a radioactive area to a minimum. Remaining in a radioactive area twice as long exposes you to twice as much radiation.

Keep your distance! The greater the distance from the radioactive source, the lower the intensity of radiation received. By doubling your distance from the radiation source, the intensity of radiation drops to \( \left( \frac{1}{2} \right)^2 \), or one-fourth of its previous value.

### Table 16.3 Properties of Radiation and Shielding Required

<table>
<thead>
<tr>
<th>Property</th>
<th>Alpha (( \alpha )) Particle</th>
<th>Beta (( \beta )) Particle</th>
<th>Gamma (( \gamma )) Ray</th>
</tr>
</thead>
<tbody>
<tr>
<td>Travel Distance in Air</td>
<td>2 to 4 cm</td>
<td>200 to 300 cm</td>
<td>500 m</td>
</tr>
<tr>
<td>Tissue Depth</td>
<td>0.05 mm</td>
<td>4 to 5 mm</td>
<td>50 cm or more</td>
</tr>
<tr>
<td>Shielding</td>
<td>Paper, clothing</td>
<td>Heavy clothing, lab coats, gloves</td>
<td>Lead, thick concrete</td>
</tr>
<tr>
<td>Typical Source</td>
<td>Radium-226</td>
<td>Carbon-14</td>
<td>Technetium-99m</td>
</tr>
</tbody>
</table>

### Questions and Problems

#### 16.1 Natural Radioactivity

**Learning Goal** Describe alpha, beta, positron, and gamma radiation.

16.1 Identify the type of particle or radiation for each of the following:
   - a. \( ^4_2 \text{He} \)
   - b. \( ^0_+ \text{e} \)
   - c. \( ^0_0 \gamma \)

16.2 Identify the type of particle or radiation for each of the following:
   - a. \( _0^- \text{e} \)
   - b. \( ^1_1 \text{H} \)
   - c. \( ^0_0 \text{n} \)

   - a. Write the atomic symbol for each isotope.
   - b. In what ways are the isotopes similar, and in what ways do they differ?

16.4 Naturally occurring iodine is iodine-127. Medically, radioactive isotopes of iodine-125 and iodine-131 are used.
   - a. Write the atomic symbol for each isotope.
   - b. In what ways are the isotopes similar, and in what ways do they differ?

16.5 Identify each of the following:
   - a. \( ^{12}_5 \text{X} \)
   - b. \( ^{13}_5 \text{X} \)
   - c. \( ^{14}_5 \text{X} \)
   - d. \( ^{35}_18 \text{X} \)
   - e. \( ^{12}_6 \text{X} \)

16.6 Identify each of the following:
   - a. \( ^{13}_5 \text{X} \)
   - b. \( ^{14}_5 \text{X} \)
   - c. \( ^{15}_5 \text{X} \)
   - d. \( ^{35}_18 \text{X} \)
   - e. \( ^{12}_6 \text{X} \)

### Applications

16.7 Write the symbol for each of the following isotopes used in nuclear medicine:
   - a. copper-64
   - b. selenium-75
   - c. sodium-24
   - d. nitrogen-15

16.8 Write the symbol for each of the following isotopes used in nuclear medicine:
   - a. indium-111
   - b. palladium-103
   - c. barium-131
   - d. rubidium-82

16.9 Supply the missing information in the following table:

<table>
<thead>
<tr>
<th>Medical Use</th>
<th>Atomic Symbol</th>
<th>Mass Number</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Heart imaging</td>
<td>201(^{81}\text{Tl})</td>
<td>81</td>
<td>55</td>
<td>27</td>
</tr>
<tr>
<td>Radiation therapy</td>
<td>60</td>
<td>27</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Abdominal scan</td>
<td>31</td>
<td>36</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hyperthyroidism</td>
<td>131(^{53}\text{I})</td>
<td>53</td>
<td>32</td>
<td>17</td>
</tr>
<tr>
<td>Leukemia treatment</td>
<td>32</td>
<td>17</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

16.10 Supply the missing information in the following table:

<table>
<thead>
<tr>
<th>Medical Use</th>
<th>Atomic Symbol</th>
<th>Mass Number</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cancer treatment</td>
<td>133(^{53}\text{Cs})</td>
<td>53</td>
<td>43</td>
<td>56</td>
</tr>
<tr>
<td>Brain scan</td>
<td>43</td>
<td>56</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Blood flow</td>
<td>141</td>
<td>58</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Bone scan</td>
<td>85</td>
<td>53</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Lung function</td>
<td>133(^{54}\text{Xe})</td>
<td>54</td>
<td>58</td>
<td>47</td>
</tr>
</tbody>
</table>

16.11 Match the type of radiation (1 to 3) with each of the following statements:
   - 1. alpha particle
   - 2. beta particle
   - 3. gamma radiation
   - a. does not penetrate skin
   - b. shielding protection includes lead or thick concrete
   - c. can be very harmful if ingested

16.12 Match the type of radiation (1 to 3) with each of the following statements:
   - 1. alpha particle
   - 2. beta particle
   - 3. gamma radiation
   - a. penetrates farthest into skin and body tissues
   - b. shielding protection includes lab coats and gloves
   - c. travels only a short distance in air
**16.2 Nuclear Reactions**

**LEARNING GOAL** Write a balanced nuclear equation showing mass numbers and atomic numbers for radioactive decay.

In a process called radioactive decay, a nucleus spontaneously breaks down by emitting radiation. This process is shown by writing a nuclear equation with the atomic symbols of the original radioactive nucleus on the left, an arrow, and the new nucleus and the type of radiation emitted on the right.

Radioactive nucleus $\rightarrow$ new nucleus + radiation ($\alpha$, $\beta$, $\beta^+$, $\gamma$)

In a nuclear equation, the sum of the mass numbers and the sum of the atomic numbers on one side of the arrow must equal the sum of the mass numbers and the sum of the atomic numbers on the other side.

**Alpha Decay**

An unstable nucleus may emit an alpha particle, which consists of two protons and two neutrons. Thus, the mass number of the radioactive nucleus decreases by 4, and its atomic number decreases by 2. For example, when uranium-238 emits an alpha particle, the new nucleus that forms has a mass number of 234. Compared to uranium with 92 protons, the new nucleus has 90 protons, which is thorium.

![Diagram of radioactive uranium nucleus emitting an alpha particle to become a thorium-234 nucleus, showing mass numbers and atomic numbers](image)

We can look at writing a balanced nuclear equation for americium-241, which undergoes alpha decay as shown in Sample Problem 16.2.

**SAMPLE PROBLEM 16.2 Writing an Equation for Alpha Decay**

Smoke detectors that are used in homes and apartments contain americium-241, which undergoes alpha decay. When alpha particles collide with air molecules, charged particles are produced that generate an electrical current. If smoke particles enter the detector, they interfere with the formation of charged particles in the air, and the electrical current is interrupted. This causes the alarm to sound and warns the occupants of the danger of fire. Write the balanced nuclear equation for the decay of americium-241.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th><strong>Analyse the problem</strong></th>
<th><strong>Given</strong></th>
<th><strong>Need</strong></th>
<th><strong>Connect</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>Am-241, alpha decay</td>
<td>balanced nuclear equation</td>
<td>mass number, atomic number of new nucleus</td>
<td></td>
</tr>
</tbody>
</table>
**Chemistry Link to Health**

**Radon in Our Homes**

The presence of radon gas has become a much publicized environmental and health issue because of the radiation danger it poses. Radioactive isotopes such as radium-226 are naturally present in many types of rocks and soils. Radium-226 emits an alpha particle and is converted into radon gas, which diffuses out of the rocks and soil.

\[
^{226}_{88}\text{Ra} \rightarrow ^{222}_{86}\text{Rn} + ^{4}_{2}\text{He}
\]

Outdoors, radon gas poses little danger because it disperses in the air. However, if the radioactive source is under a house or building, the radon gas can enter the house through cracks in the foundation or other openings. Those who live or work there may inhale the radon. Inside the lungs, radon-222 emits alpha particles to form polonium-218, which is known to cause lung cancer.

\[
^{222}_{86}\text{Rn} \rightarrow ^{218}_{84}\text{Po} + ^{4}_{2}\text{He}
\]

The U.S. Environmental Protection Agency (EPA) estimates that radon causes about 20000 lung cancer deaths in one year. The EPA recommends that the maximum level of radon not exceed 4 picocuries (pCi) per liter of air in a home. One picocurie (pCi) is equal to \(10^{-12}\) curies (Ci); curies are described in section 16.3. The EPA estimates that more than 6 million homes have radon levels that exceed this maximum.

A radon gas detector is used to determine radon levels in buildings.

**Beta Decay**

The formation of a beta particle is the result of the breakdown of a neutron into a proton and an electron (beta particle). Because the proton remains in the nucleus, the number of protons increases by one, whereas the number of neutrons decreases by one. Thus, in a

**Chemistry Link to Health**

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nuclear equation for beta decay, the mass number of the radioactive nucleus and the mass number of the new nucleus are the same. However, the atomic number of the new nucleus increases by one, which makes it a nucleus of a different element (transmutation). For example, the beta decay of a carbon-14 nucleus produces a nitrogen-14 nucleus.

What happens to a C\textsubscript{14} nucleus when a beta particle is emitted?

\begin{align*}
\text{Radioactive carbon nucleus} & \quad \text{Beta particle} \\
\begin{array}{c}
\text{Radiation} \\
\text{Neutron} \\
\text{Proton}
\end{array} & \quad \begin{array}{c}
\text{Stable nitrogen-14 nucleus} \\
\text{Neutron} \\
\text{Proton}
\end{array}
\end{align*}

In the nuclear equation for beta decay, the mass number of the new nucleus remains the same and its atomic number increases by 1.

SAMPLE PROBLEM 16.3 Writing a Nuclear Equation for Beta Decay

The radioactive isotope yttrium-90, a beta emitter, is used in cancer treatment and as a colloidal injection into large joints to relieve the pain caused by arthritis. Write the balanced nuclear equation for the beta decay of yttrium-90.

\begin{align*}
\text{TRY IT FIRST} & \\
\text{SOLUTION} & \\
\text{ANALYZE THE PROBLEM} & \begin{array}{|c|c|c|}
\hline
\text{Given} & \text{Need} & \text{Connect} \\
\hline
\text{Y-90, beta decay} & \text{balanced nuclear equation} & \text{mass number, atomic number of new nucleus} \\
\hline
\end{array}
\end{align*}

\text{STEP 1 Write the incomplete nuclear equation.}

\[
\begin{align*}
{\text{\text{\textsubscript{90}Y}}} & \quad \rightarrow \quad ? + \text{\textsubscript{-1}e} \\
{\text{Y}} & \quad \rightarrow \quad ? + \text{\textsubscript{-1}e}
\end{align*}
\]

\text{STEP 2 Determine the missing mass number. In the equation, the mass number of the yttrium, 90, is equal to the sum of the mass numbers of the new nucleus and the beta particle.}

\[
\begin{align*}
90 & = ? + 0 \\
90 - 0 & = ? \\
90 - 0 & = 90 \text{ (mass number of new nucleus)}
\end{align*}
\]

\text{STEP 3 Determine the missing atomic number. The atomic number of yttrium, 39, must equal the sum of the atomic numbers of the new nucleus and the beta particle.}

\[
\begin{align*}
39 & = ? - 1 \\
39 + 1 & = ? \\
39 + 1 & = 40 \text{ (atomic number of new nucleus)}
\end{align*}
\]

\text{STEP 4 Determine the symbol of the new nucleus. On the periodic table, the element that has atomic number 40 is zirconium, Zr. The nucleus of this isotope of Zr is written as \text{\textsubscript{90}Zr}.}

A radioisotope is injected into the joint to relieve the pain caused by arthritis.
STEP 5 Complete the nuclear equation.
\[ ^{90}_{39}\text{Y} \rightarrow ^{90}_{40}\text{Zr} + ^0_{-1}\text{e} \]

STUDY CHECK 16.3
Write the balanced nuclear equation for the beta decay of chromium-51.

ANSWER
\[ ^{51}_{24}\text{Cr} \rightarrow ^{51}_{25}\text{Mn} + ^0_{-1}\text{e} \]

Positron Emission
In positron emission, a proton in an unstable nucleus is converted to a neutron and a positron. The neutron remains in the nucleus, but the positron is emitted from the nucleus. In a nuclear equation for positron emission, the mass number of the radioactive nucleus and the mass number of the new nucleus are the same. However, the atomic number of the new nucleus decreases by one, indicating a change of one element into another. For example, an aluminum-24 nucleus undergoes positron emission to produce a magnesium-24 nucleus. The atomic number of magnesium (12) and the charge of the positron (1+) give the atomic number of aluminum (13).

\[ ^{24}_{13}\text{Al} \rightarrow ^{24}_{12}\text{Mg} + ^0_{+1}\text{e} \]

SAMPLE PROBLEM 16.4 Writing a Nuclear Equation for Positron Emission
Write the balanced nuclear equation for manganese-49, which decays by emitting a positron.

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Mn-49, positron emission</td>
<td>balanced nuclear equation</td>
<td>mass number, atomic number of new nucleus</td>
</tr>
</tbody>
</table>

STEP 1 Write the incomplete nuclear equation.
\[ ^{49}_{25}\text{Mn} \rightarrow ? + ^0_{+1}\text{e} \]

STEP 2 Determine the missing mass number. In the equation, the mass number of the manganese, 49, is equal to the sum of the mass numbers of the new nucleus and the positron.

\[ 49 = ? + 0 \]
\[ 49 - 0 = ? \]
\[ 49 - 0 = 49 \text{ (mass number of new nucleus)} \]

STEP 3 Determine the missing atomic number. The atomic number of manganese, 25, must equal the sum of the atomic numbers of the new nucleus and the positron.

\[ 25 = ? + 1 \]
\[ 25 - 1 = ? \]
\[ 25 - 1 = 24 \text{ (atomic number of new nucleus)} \]

STEP 4 Determine the symbol of the new nucleus. On the periodic table, the element that has atomic number 24 is chromium (Cr). The nucleus of this isotope of Cr is written as \(^{49}_{24}\text{Cr}\).

STEP 5 Complete the nuclear equation.
\[ ^{49}_{25}\text{Mn} \rightarrow ^{49}_{24}\text{Cr} + ^0_{+1}\text{e} \]
STUDY CHECK 16.4
Write the balanced nuclear equation for xenon-118, which undergoes positron emission.

**ANSWER**

\[ {^{118}_{54}} \text{Xe} \rightarrow {^{118}_{53}} \text{I} + {^0_0} \beta^+ \]

**Gamma Emission**

Pure gamma emitters are rare, although gamma radiation accompanies most alpha and beta radiation. In radiology, one of the most commonly used gamma emitters is technetium (Tc). The unstable isotope of technetium is written as the metastable (symbol m) isotope technetium-99m, Tc-99m, or \(^{99m}\)Tc. By emitting energy in the form of gamma rays, the nucleus becomes more stable.

\[ {^{99m}_{43}} \text{Tc} \rightarrow {^{99}_{43}} \text{Tc} + 0\gamma \]

**FIGURE 16.2** summarizes the changes in the nucleus for alpha, beta, positron, and gamma radiation.

**Producing Radioactive Isotopes**

Today, many radioisotopes are produced in small amounts by bombarding stable, nonradioactive isotopes with high-speed particles such as alpha particles, protons, neutrons, and small nuclei. When one of these particles is absorbed, the stable nucleus is converted to a radioactive isotope and usually some type of radiation particle.

**FIGURE 16.2** When the nuclei of alpha, beta, positron, and gamma emitters emit radiation, new, more stable nuclei are produced.

What changes occur in the number of protons and neutrons when an alpha emitter gives off radiation?

---

When nonradioactive B-10 is bombarded by an alpha particle, the products are radioactive N-13 and a neutron.

All elements that have an atomic number greater than 92 have been produced by bombardment. Most have been produced in only small amounts and exist for only a short time, making it difficult to study their properties. For example, when californium-249 is bombarded with nitrogen-15, the radioactive element dubnium-260 and four neutrons are produced.

\[ {^{15}_{7}} \text{N} + {^{249}_{98}} \text{Cf} \rightarrow {^{260}_{105}} \text{Db} + 4\, {^0_0} \text{n} \]

Technetium-99m is a radioisotope used in nuclear medicine for several diagnostic procedures, including the detection of brain tumors and examinations of the liver and spleen. The source of technetium-99m is molybdenum-99, which is produced in a nuclear reactor by neutron bombardment of molybdenum-98.

\[ {^1_0} \beta^+ + {^{98}_{42}} \text{Mo} \rightarrow {^{99}_{42}} \text{Mo} \]
Many radiology laboratories have small generators containing molybdenum-99, which decays to the technetium-99m radioisotope.

\[ ^{99}_{42}\text{Mo} \rightarrow ^{99}_{43}\text{Tc} + ^{0}_{-1}\text{e} \]

The technetium-99m radioisotope decays by emitting gamma rays. Gamma emission is desirable for diagnostic work because the gamma rays pass through the body to the detection equipment.

\[ ^{99}_{43}\text{Tc} \rightarrow ^{99}_{43}\text{Tc} + ^{0}_{0}\gamma \]

### SAMPLE PROBLEM 16.5 Writing an Equation for an Isotope Produced by Bombardment

Write the balanced nuclear equation for the bombardment of nickel-58 by a proton, \(^{1}_{1}\text{H}\), which produces a radioactive isotope and an alpha particle.

#### TRY IT FIRST

**SOLUTION**

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>proton bombardment</td>
<td>balanced nuclear equation</td>
<td>mass number, atomic number of new nucleus</td>
<td></td>
</tr>
<tr>
<td>of Ni-58</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Write the incomplete nuclear equation.

\[ ^{1}_{1}\text{H} + ^{58}_{28}\text{Ni} \rightarrow ? + ^{4}_{2}\text{He} \]

**STEP 2** Determine the missing mass number. In the equation, the sum of the mass numbers of the proton, 1, and the nickel, 58, must equal the sum of the mass numbers of the new nucleus and the alpha particle.

\[ 1 + 58 = ? + 4 \]
\[ 59 - 4 = ? \]
\[ 59 - 4 = 55 \text{ (mass number of new nucleus)} \]

**STEP 3** Determine the missing atomic number. The sum of the atomic numbers of the proton, 1, and nickel, 28, must equal the sum of the atomic numbers of the new nucleus and the alpha particle.

\[ 1 + 28 = ? + 2 \]
\[ 29 - 2 = ? \]
\[ 29 - 2 = 27 \text{ (atomic number of new nucleus)} \]

**STEP 4** Determine the symbol of the new nucleus. On the periodic table, the element that has atomic number 27 is cobalt, Co. The nucleus of this isotope of Co is written as \(^{55}_{27}\text{Co}\).

**STEP 5** Complete the nuclear equation.

\[ ^{1}_{1}\text{H} + ^{58}_{28}\text{Ni} \rightarrow ^{55}_{27}\text{Co} + ^{4}_{2}\text{He} \]

### STUDY CHECK 16.5

The first radioactive isotope was produced in 1934 by the bombardment of aluminum-27 by an alpha particle to produce a radioactive isotope and one neutron. Write the balanced nuclear equation for this bombardment.

**ANSWER**

\[ ^{4}_{2}\text{He} + ^{27}_{13}\text{Al} \rightarrow ^{30}_{15}\text{P} + ^{1}_{0}\text{n} \]
16.13 Write a balanced nuclear equation for the alpha decay of each of the following radioactive isotopes:

- a. \(^{208}_{82}\)Po
- b. \(^{232}_{90}\)Th
- c. \(^{251}_{102}\)No
- d. radon-220

16.14 Write a balanced nuclear equation for the alpha decay of each of the following radioactive isotopes:

- a. curium-243
- b. \(^{253}_{95}\)Es
- c. \(^{251}_{99}\)Cf
- d. iron-60

16.15 Write a balanced nuclear equation for the beta decay of each of the following radioactive isotopes:

- a. \(^{68}_{11}\)Na
- b. \(^{20}_{8}\)O
- c. strontium-92
- d. iron-60

16.16 Write a balanced nuclear equation for the beta decay of each of the following radioactive isotopes:

- a. \(^{53}_{21}\)K
- b. iron-59
- c. potassium-42
- d. \(^{14}_{55}\)Ba

16.17 Write a balanced nuclear equation for the positron emission of each of the following radioactive isotopes:

- a. silicon-26
- b. cobalt-54
- c. \(^{87}_{37}\)Rb
- d. \(^{93}_{43}\)Rh

16.18 Write a balanced nuclear equation for the positron emission of each of the following radioactive isotopes:

- a. boron-8
- b. \(^{15}_{8}\)O
- c. \(^{40}_{19}\)K
- d. nitrogen-13

16.19 Complete each of the following nuclear equations and describe the type of radiation:

- a. \(^{28}_{13}\)Al \(\rightarrow\) ? + \(\frac{0}{0}\)e
- b. \(^{180}_{75}\)Ta \(\rightarrow\) \(^{186}_{75}\)Ta + ?
- c. \(^{64}_{29}\)Cu \(\rightarrow\) \(^{66}_{30}\)Zn + ?
- d. ? \(\rightarrow\) \(^{234}_{91}\)Np + \(\frac{1}{2}\)He
- e. \(^{88}_{38}\)Hg \(\rightarrow\) ? + \(\frac{1}{2}\)e

16.20 Complete each of the following nuclear equations and describe the type of radiation:

- a. \(^{11}_{5}\)C \(\rightarrow\) \(^{11}_{6}\)B + ?
- b. \(^{35}_{16}\)S \(\rightarrow\) ? + \(\frac{0}{0}\)e
- c. ? \(\rightarrow\) \(^{90}_{37}\)Y + \(\frac{1}{0}\)e
- d. \(^{210}_{82}\)Bi \(\rightarrow\) ? + \(\frac{3}{2}\)He
- e. ? \(\rightarrow\) \(^{89}_{39}\)Y + \(\frac{1}{0}\)e

16.21 Complete each of the following bombardment reactions:

- a. \(\frac{1}{0}\)n + \(^{3}\)Be \(\rightarrow\) ?
- b. \(\frac{1}{0}\)n + \(^{135}_{53}\)Te \(\rightarrow\) ? + \(\frac{0}{0}\)e
- c. \(\frac{1}{0}\)n + ? \(\rightarrow\) \(^{24}_{11}\)Na + \(\frac{1}{2}\)He
- d. \(^{4}\)He + \(^{1}\)H \(\rightarrow\) ? + \(\frac{2}{1}\)H

16.22 Complete each of the following bombardment reactions:

- a. ? + \(^{10}_{6}\)Ar \(\rightarrow\) \(^{11}_{7}\)K + \(\frac{1}{2}\)He
- b. \(\frac{1}{0}\)n + \(^{238}_{92}\)U \(\rightarrow\) ?
- c. \(\frac{1}{0}\)n + ? \(\rightarrow\) \(^{13}_{6}\)C + \(\frac{1}{2}\)He
- d. ? + \(^{26}_{12}\)Ni \(\rightarrow\) \(^{27}_{12}\)Rg + \(\frac{1}{0}\)n

16.3 Radiation Measurement

LEARNING GOAL Describe the detection and measurement of radiation.

One of the most common instruments for detecting beta and gamma radiation is the Geiger counter. It consists of a metal tube filled with a gas such as argon. When radiation enters a window on the end of the tube, it forms charged particles in the gas, which produce an electrical current. Each burst of current is amplified to give a click and a reading on a meter.

\[ \text{Ar} + \text{radiation} \rightarrow \text{Ar}^+ + e^- \]

Measuring Radiation

Radiation is measured in several different ways. When a radiology laboratory obtains a radioisotope, the activity of the sample is measured in terms of the number of nuclear disintegrations per second. The curie (Ci), the original unit of activity, was defined as the
number of disintegrations that occur in 1 s for 1 g of radium, which is equal to $3.7 \times 10^{10}$ disintegrations/s. The unit was named for the Polish scientist Marie Curie, who along with her husband, Pierre, discovered the radioactive elements radium and polonium. The SI unit of radiation activity is the becquerel (Bq), which is 1 disintegration/s.

The rad (radiation absorbed dose) is a unit that measures the amount of radiation absorbed by a gram of material such as body tissue. The SI unit for absorbed dose is the gray (Gy), which is defined as the joules of energy absorbed by 1 kg of body tissue. The gray is equal to 100 rad.

The rem (radiation equivalent in humans) is a unit that measures the biological effects of different kinds of radiation. Although alpha particles do not penetrate the skin, if they should enter the body by some other route, they can cause extensive damage within a short distance in tissue. High-energy radiation, such as beta particles, high-energy protons, and neutrons that travel into tissue, causes more damage. Gamma rays are damaging because they travel a long way through body tissue.

To determine the equivalent dose or rem dose, the absorbed dose (rad) is multiplied by a factor that adjusts for biological damage caused by a particular form of radiation. For beta and gamma radiation the factor is 1, so the biological damage in rems is the same as the absorbed radiation (rad). For high-energy protons and neutrons, the factor is about 10, and for alpha particles it is 20.

$$\text{Biological damage (rem)} = \text{Absorbed dose (rad)} \times \text{Factor}$$

Often, the measurement for an equivalent dose will be in units of millirems (mrem). One rem is equal to 1000 mrem. The SI unit is the sievert (Sv). One sievert is equal to 100 rem.

**TABLE 16.4** summarizes the units used to measure radiation.

---

**CHEMISTRY LINK TO HEALTH**

**Radiation and Food**

Foodborne illnesses caused by pathogenic bacteria such as *Salmonella*, *Listeria*, and *Escherichia coli* have become a major health concern in the United States. *E. coli* has been responsible for outbreaks of illness from contaminated ground beef, fruit juices, lettuce, and alfalfa sprouts.

The U.S. Food and Drug Administration (FDA) has approved the use of 0.3 kGy to 1 kGy of radiation produced by cobalt-60 or cesium-137 for the treatment of foods. The irradiation technology is much like that used to sterilize medical supplies. Cobalt pellets are placed in stainless steel tubes, which are arranged in racks. When food moves through the series of racks, the gamma rays pass through the food and kill the bacteria.

It is important for consumers to understand that when food is irradiated, it never comes in contact with the radioactive source. The gamma rays pass through the food to kill bacteria, but that does not make the food radioactive. The radiation kills bacteria because it stops their ability to divide and grow. We cook or heat food thoroughly for the same purpose. Radiation, as well as heat, has little effect on the food itself because its cells are no longer dividing or growing. Thus irradiated food is not harmed, although a small amount of vitamin B$_1$ and C may be lost.

Currently, tomatoes, blueberries, strawberries, and mushrooms are being irradiated to allow them to be harvested when completely ripe and extend their shelf life (see **FIGURE 16.3**). The FDA has also approved the irradiation of pork, poultry, and beef to decrease potential infections and to extend shelf life. Currently, irradiated vegetable and meat products are available in retail markets in more than 40 countries. In the United States, irradiated foods such as tropical fruits, spinach, and ground meats are found in some stores. *Apollo 17* astronauts ate irradiated foods on the Moon, and some U.S. hospitals and nursing homes now use irradiated poultry to reduce the possibility of salmonella infections among residents. The extended shelf life of irradiated food also makes it useful for campers and military personnel. Soon, consumers concerned about food safety will have a choice of irradiated meats, fruits, and vegetables at the market.

---

**FIGURE 16.3** (a) The FDA requires this symbol to appear on irradiated retail foods. (b) After two weeks, the irradiated strawberries on the right show no spoilage. Mold is growing on the nonirradiated ones on the left.

**Why are irradiated foods used on spaceships and in nursing homes?**
People who work in radiology laboratories wear dosimeters attached to their clothing to determine any exposure to radiation such as X-rays, gamma rays, or beta particles. A dosimeter can be thermoluminescent (TLD), optically stimulated luminescence (OSL), or electronic personal (EPD). Dosimeters provide real time radiation levels measured by monitors in the work area.

**Sample Problem 16.6 Radiation Measurement**

One treatment for bone pain involves intravenous administration of the radioisotope phosphorus-32, which is incorporated into bone. A typical dose of 7 mCi can produce up to 450 rad in the bone. What is the difference between the units of mCi and rad?

**TRY IT FIRST**

**Solution**

The millicuries (mCi) indicate the activity of the P-32 in terms of nuclei that break down in 1 s. The radiation absorbed dose (rad) is a measure of amount of radiation absorbed by the bone.

**Study Check 16.6**

For Sample Problem 16.6, what is the absorbed dose of radiation in grays (Gy)?

**Answer**

4.5 Gy

**Exposure to Radiation**

Every day, we are exposed to low levels of radiation from naturally occurring radioactive isotopes in the buildings where we live and work, in our food and water, and in the air we breathe. For example, potassium-40, a naturally occurring radioactive isotope, is present in any potassium-containing food. Other naturally occurring radioisotopes in air and food are carbon-14, radon-222, strontium-90, and iodine-131. The average person in the United States is exposed to about 360 mrem of radiation annually. Medical sources of radiation, including dental, hip, spine, and chest X-rays and mammograms, add to our radiation exposure. Table 16.5 lists some common sources of radiation.

Another source of background radiation is cosmic radiation produced in space by the Sun. People who live at high altitudes or travel by airplane receive a greater amount of cosmic radiation because there are fewer molecules in the atmosphere to absorb the radiation. For example, a person living in Denver receives about twice the cosmic radiation as a person living in Los Angeles. A person living close to a nuclear power plant normally does not receive much additional radiation, perhaps 0.1 mrem in 1 yr. (One rem equals 1000 mrem.) However, in the accident at the Chernobyl nuclear power plant in 1986 in Ukraine, it is estimated that people in a nearby town received as much as 1 rem/h.

**Table 16.4 Units of Radiation Measurement**

<table>
<thead>
<tr>
<th>Measurement</th>
<th>Common Unit</th>
<th>SI Unit</th>
<th>Relationship</th>
</tr>
</thead>
<tbody>
<tr>
<td>Activity</td>
<td>curie (Ci)</td>
<td>becquerel (Bq)</td>
<td>$1 \text{ Ci} = 3.7 \times 10^{10} \text{ Bq}$</td>
</tr>
<tr>
<td></td>
<td>1 Ci = $3.7 \times 10^{10}$ disintegrations/s</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>rad</td>
<td>gray (Gy)</td>
<td>1 Gy = 1 J/kg of tissue</td>
</tr>
<tr>
<td>Absorbed Dose</td>
<td></td>
<td></td>
<td>1 Gy = 100 rad</td>
</tr>
<tr>
<td>Biological Damage</td>
<td>rem</td>
<td>sievert (Sv)</td>
<td>1 Sv = 100 rem</td>
</tr>
</tbody>
</table>

**Table 16.5 Average Annual Radiation Received by a Person in the United States**

<table>
<thead>
<tr>
<th>Source</th>
<th>Dose (mrem)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Natural</td>
<td></td>
</tr>
<tr>
<td>Ground</td>
<td>20</td>
</tr>
<tr>
<td>Air, water, food</td>
<td>30</td>
</tr>
<tr>
<td>Cosmic rays</td>
<td>40</td>
</tr>
<tr>
<td>Wood, concrete, brick</td>
<td>50</td>
</tr>
<tr>
<td>Medical</td>
<td></td>
</tr>
<tr>
<td>Chest X-ray</td>
<td>20</td>
</tr>
<tr>
<td>Dental X-ray</td>
<td>20</td>
</tr>
<tr>
<td>Mammogram</td>
<td>40</td>
</tr>
<tr>
<td>Hip X-ray</td>
<td>60</td>
</tr>
<tr>
<td>Lumbar spine X-ray</td>
<td>70</td>
</tr>
<tr>
<td>Upper gastrointestinal tract X-ray</td>
<td>200</td>
</tr>
<tr>
<td>Other</td>
<td></td>
</tr>
<tr>
<td>Nuclear power plants</td>
<td>0.1</td>
</tr>
<tr>
<td>Television</td>
<td>20</td>
</tr>
<tr>
<td>Air travel</td>
<td>10</td>
</tr>
<tr>
<td>Radon</td>
<td>200*</td>
</tr>
</tbody>
</table>

*Varies widely.
**Radiation Sickness**

The larger the dose of radiation received at one time, the greater the effect on the body. Exposure to radiation of less than 25 rem usually cannot be detected. Whole-body exposure of 100 rem produces a temporary decrease in the number of white blood cells. If the exposure to radiation is greater than 100 rem, a person may suffer the symptoms of radiation sickness: nausea, vomiting, fatigue, and a reduction in white-cell count. A whole-body dosage greater than 300 rem can decrease the white-cell count to zero. The person suffers diarrhea, hair loss, and infection. Exposure to radiation of 500 rem is expected to cause death in 50% of the people receiving that dose. This amount of radiation to the whole body is called the *lethal dose for one-half the population*, or the LD$_{50}$. The LD$_{50}$ varies for different life forms, as Table 16.6 shows. Whole-body radiation of 600 rem or greater would be fatal to all humans within a few weeks.

**Table 16.6 Lethal Doses of Radiation for Some Life Forms**

<table>
<thead>
<tr>
<th>Life Form</th>
<th>LD$_{50}$ (rem)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Insect</td>
<td>100 000</td>
</tr>
<tr>
<td>Bacterium</td>
<td>50 000</td>
</tr>
<tr>
<td>Rat</td>
<td>800</td>
</tr>
<tr>
<td>Human</td>
<td>500</td>
</tr>
<tr>
<td>Dog</td>
<td>300</td>
</tr>
</tbody>
</table>

**Questions and Problems**

**16.3 Radiation Measurement**

**LEARNING GOAL** Describe the detection and measurement of radiation.

**16.23** Match each property (1 to 3) with its unit of measurement.

1. activity  
2. absorbed dose  
3. biological damage

a. rad  
b. mrem  
c. mCi  
d. Gy

**16.24** Match each property (1 to 3) with its unit of measurement.

1. activity  
2. absorbed dose  
3. biological damage

a. mrad  
b. gray  
c. becquerel  
d. Sv

**Applications**

16.25 Two technicians in a nuclear laboratory were accidentally exposed to radiation. If one was exposed to 8 mGy and the other to 5 rad, which technician received more radiation?

16.26 Two samples of a radioisotope were spilled in a nuclear laboratory. The activity of one sample was 8 kBq and the other 15 mCi. Which sample produced the higher amount of radiation?

16.27 a. The recommended dosage of iodine-131 is 4.20 µCi/kg of body mass. How many microcuries of iodine-131 are needed for a 70.0-kg person with hyperthyroidism?  
   b. A person receives 50 rad of gamma radiation. What is that amount in grays?

16.28 a. The dosage of technetium-99m for a lung scan is 20. μCi/kg of body mass. How many millicuries of technetium-99m should be given to a 50.0-kg person (1 mCi = 1000 μCi)?  
   b. Suppose a person absorbed 50 mrad of alpha radiation. What would be the equivalent dose in millirems?

**16.4 Half-Life of a Radioisotope**

**LEARNING GOAL** Given the half-life of a radioisotope, calculate the amount of radioisotope remaining after one or more half-lives.

The *half-life* of a radioisotope is the amount of time it takes for one-half of a sample to decay. For example, $^{131}\text{I}$ has a half-life of 8.0 days. As $^{131}\text{I}$ decays, it produces the nonradioactive isotope $^{131}\text{Xe}$ and a beta particle.

$$^{131}\text{I} \rightarrow ^{131}\text{Xe} + _{-1}^0\text{e}$$

Suppose we have a sample that initially contains 20. mg of $^{131}\text{I}$. In 8.0 days, one-half (10. mg) of all the I-131 nuclei in the sample will decay, which leaves 10. mg of I-131. After 16 days (two half-lives), 5.0 mg of the remaining I-131 decays, which leaves 5.0 mg of I-131. After 24 days (three half-lives), 2.5 mg of the remaining I-131 decays, which leaves 2.5 mg of I-131 nuclei still capable of producing radiation.
SA M P LE P ROBLEM 16.7 Using Half-Lives of a Radioisotope

Phosphorus-32, a radioisotope used in the treatment of leukemia, has a half-life of 14.3 days. If a sample contains 8.0 mg of phosphorus-32, how many milligrams of phosphorus-32 remain after 42.9 days?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>GIVEN</th>
<th>NEED</th>
<th>CONNECT</th>
</tr>
</thead>
<tbody>
<tr>
<td>8.0 mg of P-32, 42.9 days</td>
<td>milligrams of</td>
<td>number of</td>
</tr>
<tr>
<td>elapsed, half-life = 14.3 days</td>
<td>P-32 remaining</td>
<td>half-lives</td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to calculate the unknown quantity.

$$\text{days} \times \text{Half-life} = \text{number of half-lives}$$

$$\text{milligrams of } ^{32}\text{P} \times \text{Number of half-lives} = \text{milligrams of } ^{32}\text{P remaining}$$

**STEP 3** Write the half-life equality and conversion factors.

1 half-life = 14.3 days

14.3 days and 1 half-life

1 half-life = 14.3 days
**STEP 4** Set up the problem to calculate the needed quantity. First, we determine the number of half-lives in the amount of time that has elapsed.

\[
\text{number of half-lives} = \frac{42.9 \text{ days}}{14.3 \text{ days}} \times 1 \text{ half-life} = 3.00 \text{ half-lives}
\]

Now we can determine how much of the sample decays in three half-lives and how many grams of the phosphorus remain.

\[
8.0 \text{ mg of } ^{32}\text{P} \quad \xrightarrow{1\text{ half-life}} \quad 4.0 \text{ mg of } ^{32}\text{P} \quad \xrightarrow{2\text{ half-lives}} \quad 2.0 \text{ mg of } ^{32}\text{P} \quad \xrightarrow{3\text{ half-lives}} \quad 1.0 \text{ mg of } ^{32}\text{P}
\]

**STUDY CHECK 16.7**

Iron-59 has a half-life of 44 days. If a nuclear laboratory receives a sample of 8.0 μg of iron-59, how many micrograms of iron-59 are still active after 176 days?

**ANSWER**

0.50 μg of iron-59

<table>
<thead>
<tr>
<th>Table 16.7: Half-Lives of Some Radioisotopes</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Element</strong></td>
</tr>
<tr>
<td>Naturally Occurring Radioisotopes</td>
</tr>
<tr>
<td>Carbon-14</td>
</tr>
<tr>
<td>Potassium-40</td>
</tr>
<tr>
<td>Radium-226</td>
</tr>
<tr>
<td>Strontium-90</td>
</tr>
<tr>
<td>Uranium-238</td>
</tr>
<tr>
<td>Some Medical Radioisotopes</td>
</tr>
<tr>
<td>Carbon-11</td>
</tr>
<tr>
<td>Chromium-51</td>
</tr>
<tr>
<td>Iodine-131</td>
</tr>
<tr>
<td>Oxygen-15</td>
</tr>
<tr>
<td>Iron-59</td>
</tr>
<tr>
<td>Radon-222</td>
</tr>
<tr>
<td>Technetium-99m</td>
</tr>
</tbody>
</table>

Naturally occurring isotopes of the elements usually have long half-lives, as shown in Table 16.7. They disintegrate slowly and produce radiation over a long period of time, even hundreds or millions of years. In contrast, the radioisotopes used in nuclear medicine have much shorter half-lives. They disintegrate rapidly and produce almost all their radiation in a short period of time. For example, technetium-99m emits half of its radiation in the first 6 h. This means that a small amount of the radioisotope given to a patient is essentially gone within two days. The decay products of technetium-99m are totally eliminated by the body.
Sample Problem 16.8 Dating Using Half-Lives

Carbon material in the bones of humans and animals assimilates carbon until death. Using radiocarbon dating, the number of half-lives of carbon-14 from a bone sample determines the age of the bone. Suppose a sample is obtained from a prehistoric animal and used for radiocarbon dating. We can calculate the age of the bone or the years elapsed since the animal died by using the half-life of carbon-14, which is 5730 yr. A bone sample from the skeleton of a prehistoric animal has 25% of the activity of C-14 found in a living animal. How many years ago did the prehistoric animal die?

**TRY IT FIRST**

**SOLUTION**

**STEP 1** State the given and needed quantities.

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 half-life of C-14 = 5730 yr</td>
<td>years elapsed</td>
<td>number of half-lives</td>
<td></td>
</tr>
<tr>
<td>25% of initial C-14 activity</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 2** Write a plan to calculate the unknown quantity.

Activity: $100\%$ (initial) → $50\%$ (1.0 half-life) → $25\%$ (2.0 half-lives)
STEP 3 Write the half-life equality and conversion factors.

\[
\text{1 half-life} = \frac{5730 \text{ yr}}{1 \text{ half-life}} \quad \text{and} \quad \frac{1 \text{ half-life}}{5730 \text{ yr}}
\]

STEP 4 Set up the problem to calculate the needed quantity.

\[
\text{Years elapsed} = 2.0 \text{ half-lives} \times \frac{5730 \text{ yr}}{1 \text{ half-life}} = 11000 \text{ yr}
\]

We would estimate that the animal died 11000 yr ago.

STUDY CHECK 16.8

Suppose that a piece of wood found in a tomb had one-eighth of its original carbon-14 activity. About how many years ago was the wood part of a living tree?

ANSWER

17000 yr

QUESTIONS AND PROBLEMS

16.4 Half-Life of a Radioisotope

LEARNING GOAL Given the half-life of a radioisotope, calculate the amount of radioisotope remaining after one or more half-lives.

16.29 For each of the following, indicate if the number of half-lives elapsed is:
1. one half-life
2. two half-lives
3. three half-lives
   a. a sample of Pd-103 with a half-life of 17 days after 34 days
   b. a sample of C-11 with a half-life of 20 min after 20 min
   c. a sample of At-211 with a half-life of 7 h after 21 h

16.30 For each of the following, indicate if the number of half-lives elapsed is:
1. one half-life
2. two half-lives
3. three half-lives
   a. a sample of Ce-141 with a half-life of 32.5 days after 32.5 days
   b. a sample of F-18 with a half-life of 110 min after 330 min
   c. a sample of Au-198 with a half-life of 2.7 days after 5.4 days

Applications

16.31 Technetium-99m is an ideal radioisotope for scanning organs because it has a half-life of 6.0 h and is a pure gamma emitter. Suppose that 80.0 mg were prepared in the technetium generator this morning. How many milligrams of technetium-99m would remain after each of the following intervals?
   a. one half-life
   b. two half-lives
   c. 18 h
   d. 24 h

16.32 A sample of sodium-24 with an activity of 12 mCi is used to study the rate of blood flow in the circulatory system. If sodium-24 has a half-life of 15 h, what is the activity after each of the following intervals?
   a. one half-life
   b. 30 h
   c. three half-lives
   d. 2.5 days

16.33 Strontium-85, used for bone scans, has a half-life of 65 days.
   a. How long will it take for the radiation level of strontium-85 to drop to one-fourth of its original level?
   b. How long will it take for the radiation level of strontium-85 to drop to one-eighth of its original level?

16.34 Fluorine-18, which has a half-life of 110 min, is used in PET scans.
   a. If 100. mg of fluorine-18 is shipped at 8:00 a.m., how many milligrams of the radioisotope are still active after 110 min?
   b. If 100. mg of fluorine-18 is shipped at 8:00 a.m., how many milligrams of the radioisotope are still active when the sample arrives at the radiology laboratory at 1:30 p.m.?
16.5 Medical Applications Using Radioactivity

LEARNING GOAL  Describe the use of radioisotopes in medicine.

The first radioactive isotope was used to treat a person with leukemia at the University of California at Berkeley. In 1946, radioactive iodine was successfully used to diagnose thyroid function and to treat hyperthyroidism and thyroid cancer. Radioactive isotopes are now used to produce images of organs including the liver, spleen, thyroid gland, kidneys, brain, and heart.

To determine the condition of an organ in the body, a nuclear medicine technologist may use a radioisotope that concentrates in that organ. The cells in the body do not differentiate between a nonradioactive atom and a radioactive one, so these radioisotopes are easily incorporated. Then the radioactive atoms are detected because they emit radiation. Some radioisotopes used in nuclear medicine are listed in TABLE 16.8.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Half-Life</th>
<th>Radiation</th>
<th>Medical Application</th>
</tr>
</thead>
<tbody>
<tr>
<td>Au-198</td>
<td>2.7 days</td>
<td>Beta</td>
<td>Liver imaging; treatment of abdominal carcinoma</td>
</tr>
<tr>
<td>Ce-141</td>
<td>32.5 days</td>
<td>Beta</td>
<td>Gastrointestinal tract diagnosis; measuring blood flow to the heart</td>
</tr>
<tr>
<td>Cs-131</td>
<td>9.7 days</td>
<td>Gamma</td>
<td>Prostate brachytherapy</td>
</tr>
<tr>
<td>F-18</td>
<td>110 min</td>
<td>Positron</td>
<td>Positron emission tomography (PET)</td>
</tr>
<tr>
<td>Ga-67</td>
<td>78 h</td>
<td>Gamma</td>
<td>Abdominal imaging; tumor detection</td>
</tr>
<tr>
<td>Ga-68</td>
<td>68 min</td>
<td>Gamma</td>
<td>Detection of pancreatic cancer</td>
</tr>
<tr>
<td>I-123</td>
<td>13.2 h</td>
<td>Gamma</td>
<td>Treatment of thyroid, brain, and prostate cancer</td>
</tr>
<tr>
<td>I-131</td>
<td>8.0 days</td>
<td>Beta</td>
<td>Treatment of Graves’ disease, goiter, hyperthyroidism, thyroid and prostate cancer</td>
</tr>
<tr>
<td>Ir-192</td>
<td>74 days</td>
<td>Gamma</td>
<td>Treatment of breast and prostate cancer</td>
</tr>
<tr>
<td>P-32</td>
<td>14.3 days</td>
<td>Beta</td>
<td>Treatment of leukemia, excess red blood cells, and pancreatic cancer</td>
</tr>
<tr>
<td>Pd-103</td>
<td>17 days</td>
<td>Gamma</td>
<td>Prostate brachytherapy</td>
</tr>
<tr>
<td>Sr-85</td>
<td>65 days</td>
<td>Gamma</td>
<td>Detection of bone lesions; brain scans</td>
</tr>
<tr>
<td>Tc-99m</td>
<td>6.0 h</td>
<td>Gamma</td>
<td>Imaging of skeleton and heart muscle, brain, liver, heart, lungs, bone, spleen, kidney, and thyroid; most widely used radioisotope in nuclear medicine</td>
</tr>
<tr>
<td>Y-90</td>
<td>2.7 days</td>
<td>Beta</td>
<td>Treatment of liver cancer</td>
</tr>
</tbody>
</table>

SAMPLE PROBLEM 16.9  Medical Applications of Radioactivity

In the treatment of abdominal carcinoma, a person is treated with “seeds” of gold-198, which is a beta emitter. Write the balanced nuclear equation for the beta decay of gold-198.

TRY IT FIRST

SOLUTION

Given: gold-198, beta decay

Need: balanced nuclear equation

Connect: mass number, atomic number of new nucleus


**STEP 1** Write the incomplete nuclear equation.

\[ ^{198}_{79}\text{Au} \longrightarrow ? + ^0_1\text{e} \]

**STEP 2** Determine the missing mass number. In beta decay, the mass number, 198, does not change.

\[ ^{198}_{79}\text{Au} \longrightarrow ^{198}_{79}\text{Au} + ^0_1\text{e} \]

**STEP 3** Determine the missing atomic number. The atomic number of the new isotope increases by one.

\[ ^{198}_{79}\text{Au} \longrightarrow ^{198}_{80}\text{Hg} + ^0_1\text{e} \]

**STEP 4** Determine the symbol of the new nucleus. On the periodic table, mercury, Hg, has the atomic number of 80.

**STEP 5** Complete the nuclear equation.

\[ ^{198}_{79}\text{Au} \longrightarrow ^{198}_{80}\text{Hg} + ^0_1\text{e} \]

**STUDY CHECK 16.9**

In an experimental treatment, a person is given boron-10, which is taken up by malignant tumors. When bombarded with neutrons, boron-10 decays by emitting alpha particles that destroy the surrounding tumor cells. Write the balanced nuclear equation for the reaction for this experimental procedure.

**ANSWER**

\[ ^7_1\text{H} + ^{10}_5\text{B} \longrightarrow ^7_3\text{Li} + ^3_2\text{He} \]

**Scans with Radioisotopes**

After a person receives a radioisotope, the radiologist determines the level and location of radioactivity emitted by the radioisotope. An apparatus called a scanner is used to produce an image of the organ. The scanner moves slowly across the body above the region where the organ containing the radioisotope is located. The gamma rays emitted from the radioisotope in the organ can be used to expose a photographic plate, producing a scan of the organ. On a scan, an area of decreased or increased radiation can indicate conditions such as a disease of the organ, a tumor, a blood clot, or edema.

A common method of determining thyroid function is the use of radioactive iodine uptake. Taken orally, the radioisotope iodine-131 mixes with the iodine already present in the thyroid. Twenty-four hours later, the amount of iodine taken up by the thyroid is determined. A detection tube held up to the area of the thyroid gland detects the radiation coming from the iodine-131 that has located there (see **FIGURE 16.5**).

A person with a hyperactive thyroid will have a higher than normal level of radioactive iodine, whereas a person with a hypoactive thyroid will have lower values. If a person has hyperthyroidism, treatment is begun to lower the activity of the thyroid. One treatment involves giving a therapeutic dosage of radioactive iodine, which has a higher radiation level than the diagnostic dose. The radioactive iodine goes to the thyroid where its radiation destroys some of the thyroid cells. The thyroid produces less thyroid hormone, bringing the hyperthyroid condition under control.
**Positron Emission Tomography**

Positron emitters with short half-lives such as carbon-11, oxygen-15, nitrogen-13, and fluorine-18 are used in an imaging method called *positron emission tomography* (PET). A positron-emitting isotope such as fluorine-18 combined with substances in the body such as glucose is used to study brain function, metabolism, and blood flow.

\[
^{18}\text{F} \rightarrow ^{18}\text{O} + _{0}^1\text{e}^-
\]

As positrons are emitted, they combine with electrons to produce gamma rays that are detected by computerized equipment to create a three-dimensional image of the organ (see **Figure 16.6**).

**Computed Tomography**

Another imaging method used to scan organs such as the brain, lungs, and heart is *computed tomography* (CT). A computer monitors the absorption of 30,000 X-ray beams directed at successive layers of the target organ. Based on the densities of the tissues and fluids in the organ, the differences in absorption of the X-rays provide a series of images of the organ. This technique is successful in the identification of hemorrhages, tumors, and atrophy.

**Magnetic Resonance Imaging**

*Magnetic resonance imaging* (MRI) is a powerful imaging technique that does not involve X-ray radiation. It is the least invasive imaging method available. MRI is based on the absorption of energy when the protons in hydrogen atoms are excited by a strong magnetic field. Hydrogen atoms make up 63% of all the atoms in the body. In the hydrogen nuclei, the protons act like tiny bar magnets. With no external field, the protons have random orientations. However, when placed within a strong magnetic field, the protons align with the field. A proton aligned with the field has a lower energy than one that is aligned against the field. As the MRI scan proceeds, radiofrequency pulses of energy are applied and the hydrogen nuclei resonate at a certain frequency. Then the radio waves are quickly turned off, and the protons slowly return to their natural alignment within the magnetic field and resonate at a different frequency. They release the energy absorbed from the radio wave pulses. The difference in energy between the two states is released, which produces the electromagnetic signal that the scanner detects. These signals are sent to a computer system, where a color image of the body is generated. Because hydrogen atoms in the body are in different chemical environments, different energies are absorbed. MRI is particularly useful in obtaining images of soft tissues, which contain large amounts of hydrogen atoms in the form of water.

**CHEMISTRY LINK TO HEALTH**

**Brachytherapy**

The process called brachytherapy, or seed implantation, is an internal form of radiation therapy. The prefix *brachy* is from the Greek word for short distance. With internal radiation, a high dose of radiation is delivered to a cancerous area, whereas normal tissue sustains minimal damage. Because higher doses are used, fewer treatments of shorter duration are needed. Conventional external treatment delivers a lower dose per treatment, but requires six to eight weeks of treatment.

**Permanent Brachytherapy**

One of the most common forms of cancer in males is prostate cancer. In addition to surgery and chemotherapy, one treatment option is to place 40 or more titanium capsules, or “seeds,” in the malignant area. Each seed, which is the size of a grain of rice, contains radioactive iodine-125, palladium-103, or cesium-131, which decay by gamma emission. The radiation from the seeds destroys the cancer
by interfering with the reproduction of cancer cells with minimal damage to adjacent normal tissues. Ninety percent (90%) of the radioisotopes decay within a few months because they have short half-lives.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>I-125</th>
<th>Pd-103</th>
<th>Cs-131</th>
</tr>
</thead>
<tbody>
<tr>
<td>Radiation</td>
<td>Gamma</td>
<td>Gamma</td>
<td>Gamma</td>
</tr>
<tr>
<td>Half-Life</td>
<td>60 days</td>
<td>17 days</td>
<td>10 days</td>
</tr>
<tr>
<td>Time Required to Deliver 90% of Radiation</td>
<td>7 months</td>
<td>2 months</td>
<td>1 month</td>
</tr>
</tbody>
</table>

Almost no radiation passes out of the patient’s body. The amount of radiation received by a family member is no greater than that received on a long plane flight. Because the radioisotopes decay to products that are not radioactive, the inert titanium capsules can be left in the body.

**Temporary Brachytherapy**

In another type of treatment for prostate cancer, long needles containing iridium-192 are placed in the tumor. However, the needles are removed after 5 to 10 min, depending on the activity of the iridium isotope. Compared to permanent brachytherapy, temporary brachytherapy can deliver a higher dose of radiation over a shorter time. The procedure may be repeated in a few days.

Brachytherapy is also used following breast cancer lumpectomy. An iridium-192 isotope is inserted into the catheter implanted in the space left by the removal of the tumor. The isotope is removed after 5 to 10 min, depending on the activity of the iridium source. Radiation is delivered primarily to the tissue surrounding the cavity that contained the tumor and where the cancer is most likely to recur. The procedure is repeated twice a day for five days to give an absorbed dose of 34 Gy (3400 rad). The catheter is removed, and no radioactive material remains in the body.

In conventional external beam therapy for breast cancer, a patient is given 2 Gy once a day for six to seven weeks, which gives a total absorbed dose of about 80 Gy or 8000 rad. The external beam therapy irradiates the entire breast, including the tumor cavity.

**QUESTIONS AND PROBLEMS**

### 16.5 Medical Applications Using Radioactivity

**LEARNING GOAL** Describe the use of radioisotopes in medicine.

**Applications**

**16.35** Bone and bony structures contain calcium and phosphorus.

a. Why would the radioisotopes calcium-47 and phosphorus-32 be used in the diagnosis and treatment of bone diseases?

b. During nuclear tests, scientists were concerned that strontium-85, a radioactive product, would be harmful to the growth of bone in children. Explain.

**16.36**

a. Technetium-99m emits only gamma radiation. Why would this type of radiation be used in diagnostic imaging rather than an isotope that also emits beta or alpha radiation?

b. A person with polycythemia vera (excess production of red blood cells) receives radioactive phosphorus-32. Why would this treatment reduce the production of red blood cells in the bone marrow of the patient?

**16.37** In a diagnostic test for leukemia, a person receives 4.0 mL of a solution containing selenium-75. If the activity of the selenium-75 is 45 μCi/mL, what dose, in microcuries, does the patient receive?

**16.38** A vial contains radioactive iodine-131 with an activity of 2.0 mCi/mL. If a thyroid test requires 3.0 mCi in an “atomic cocktail,” how many milliliters are used to prepare the iodine-131 solution?

### 16.6 Nuclear Fission and Fusion

**LEARNING GOAL** Describe the processes of nuclear fission and fusion.

During the 1930s, scientists bombarding uranium-235 with neutrons discovered that the U-235 nucleus splits into two smaller nuclei and produces a great amount of energy. This was the discovery of nuclear fission. The energy generated by splitting the atom was called atomic energy. A typical equation for nuclear fission is:
If we could determine the mass of the products krypton, barium, and three neutrons, with great accuracy, we would find that their total mass is slightly less than the mass of the starting materials. The missing mass has been converted into an enormous amount of energy, consistent with the famous equation derived by Albert Einstein.

\[ E = mc^2 \]

where \( E \) is the energy released, \( m \) is the mass lost, and \( c \) is the speed of light, \( 3 \times 10^8 \) m/s. Even though the mass loss is very small, when it is multiplied by the speed of light squared, the result is a large value for the energy released. The fission of 1 g of uranium-235 produces about as much energy as the burning of 3 tons of coal.

**Chain Reaction**

Fission begins when a neutron collides with the nucleus of a uranium atom. The resulting nucleus is unstable and splits into smaller nuclei. This fission process also releases several neutrons and large amounts of gamma radiation and energy. The neutrons emitted have high energies and bombard other uranium-235 nuclei. In a chain reaction, there is a rapid increase in the number of high-energy neutrons available to react with more uranium. To sustain a nuclear chain reaction, sufficient quantities of uranium-235 must be brought together to provide a critical mass in which almost all the neutrons immediately collide with more uranium-235 nuclei. So much heat and energy build up that an atomic explosion can occur (see Figure 16.7).

**Nuclear Fusion**

In fusion, two small nuclei combine to form a larger nucleus. Mass is lost, and a tremendous amount of energy is released, even more than the energy released from nuclear fission. However, a fusion reaction requires a temperature of 100 000 000 °C to overcome the repulsion of the hydrogen nuclei and cause them to undergo fusion. Fusion reactions occur continuously in the Sun and other stars, providing us with heat and light. The huge amounts of energy produced by our sun come from the fusion of \( 6 \times 10^{11} \) kg of hydrogen every second. In a fusion reaction, isotopes of hydrogen combine to form helium and large amounts of energy.

Scientists expect less radioactive waste with shorter half-lives from fusion reactors. However, fusion is still in the experimental stage because the extremely high temperatures needed have been difficult to reach and even more difficult to maintain. Research groups around the world are attempting to develop the technology needed to make the harnessing of the fusion reaction for energy a reality in our lifetime.
In a nuclear chain reaction, the fission of each uranium-235 atom produces three neutrons that cause the nuclear fission of more and more uranium-235 atoms.

Why is the fission of uranium-235 called a chain reaction?

In a fusion reactor, high temperatures are needed to combine hydrogen atoms.
SAMPLE PROBLEM 16.10 Identifying Fission and Fusion

Classify the following as pertaining to fission, fusion, or both:

a. A large nucleus breaks apart to produce smaller nuclei.
b. Large amounts of energy are released.
c. Extremely high temperatures are needed for reaction.

TRY IT FIRST

SOLUTION

a. When a large nucleus breaks apart to produce smaller nuclei, the process is fission.
b. Large amounts of energy are generated in both the fusion and fission processes.
c. An extremely high temperature is required for fusion.

STUDY CHECK 16.10

Classify the following nuclear equation as pertaining to fission, fusion, or both:

\[ _1^3\text{H} + _2^4\text{H} \rightarrow _2^4\text{He} + _0^1\text{n} + \text{energy} \]

ANSWER

When small nuclei combine to release energy, the process is fusion.

QUESTIONS AND PROBLEMS 16.6 Nuclear Fission and Fusion

LEARNING GOAL Describe the processes of nuclear fission and fusion.

16.39 What is nuclear fission?
16.40 How does a chain reaction occur in nuclear fission?
16.41 Complete the following fission reaction:
\[ _0^1\text{n} + _{92}^{235}\text{U} \rightarrow _{36}^{90}\text{Sn} + ? + _{0}^{1}\text{n} + \text{energy} \]
16.42 In another fission reaction, uranium-235 bombarded with a neutron produces strontium-94, another small nucleus, and three neutrons. Write the balanced nuclear equation for the fission reaction.

16.43 Indicate whether each of the following is characteristic of the fission or fusion process, or both:
a. Neutrons bombard a nucleus.
b. The nuclear process occurs in the Sun.
c. A large nucleus splits into smaller nuclei.
d. Small nuclei combine to form larger nuclei.
16.44 Indicate whether each of the following is characteristic of the fission or fusion process, or both:
a. Very high temperatures are required to initiate the reaction.
b. Less radioactive waste is produced.
c. Hydrogen nuclei are the reactants.
d. Large amounts of energy are released when the nuclear reaction occurs.

CHEMISTRY LINK TO THE ENVIRONMENT Nuclear Power Plants

In a nuclear power plant, the quantity of uranium-235 is held below a critical mass, so it cannot sustain a chain reaction. The fission reactions are slowed by placing control rods, which absorb some of the fast-moving neutrons, among the uranium samples. In this way, less fission occurs, and there is a slower, controlled production of energy. The heat from the controlled fission is used to produce steam. The steam drives a generator, which produces electricity. Approximately 20% of the electrical energy produced in the United States is generated in nuclear power plants.

Although nuclear power plants help meet some of our energy needs, there are some problems associated with nuclear power. One of the most serious problems is the production of radioactive byproducts that have very long half-lives, such as plutonium-239 with a half-life of 24,000 yr. It is essential that these waste products be stored safely in a place where they do not contaminate the environment.
Follow Up

CARDIAC IMAGING USING A RADIOISOTOPE

As part of her nuclear stress test, Simone starts to walk on the treadmill. When she reaches the maximum level, Dr. Paul injects a radioactive dye with Tl-201 with an activity of 74 MBq. The radiation emitted from areas of heart is detected by a scanner and produces images of her heart muscle. A thallium stress test can determine how effectively coronary arteries provide blood to the heart. If Simone has any damage to her coronary arteries, reduced blood flow during stress would show a narrowing of an artery or a blockage. After Simone rests for 3 h, Dr. Paul injects more dye with Tl-201 and she is placed under the scanner again. A second set of images of her heart muscle at rest are taken. When Simone’s doctor reviews her scans, he assures her that she had normal blood flow to her heart muscle, both at rest and under stress.

Applications

16.45 What is the activity of the radioactive dye injection for Simone
   a. in curies?
   b. in millicuries?

16.46 If the half-life of Tl-201 is 3.0 days, what is its activity, in megabecquerels
   a. after 3.0 days?
   b. after 6.0 days?

16.47 How many days will it take until the activity of the Tl-201 in Simone’s body is one-eighth of the initial activity?

16.48 Radiation from Tl-201 to Simone’s kidneys can be 24 mGy. What is this amount of radiation in rads?
16.1 Natural Radioactivity

LEARNING GOAL: Describe alpha, beta, positron, and gamma radiation.

- Radioactive isotopes have unstable nuclei that break down (decay), spontaneously emitting alpha (α), beta (β), positron (β⁺), and gamma (γ) radiation.
- Because radiation can damage the cells in the body, proper protection must be used: shielding, limiting the time of exposure, and distance.

16.2 Nuclear Reactions

LEARNING GOAL: Write a balanced nuclear equation showing mass numbers and atomic numbers for radioactive decay.

- A balanced nuclear equation is used to represent the changes that take place in the nuclei of the reactants and products.
- The new isotopes and the type of radiation emitted can be determined from the symbols that show the mass numbers and atomic numbers of the isotopes in the nuclear equation.
- A radioisotope is produced artificially when a nonradioactive isotope is bombarded by a small particle.

16.3 Radiation Measurement

LEARNING GOAL: Describe the detection and measurement of radiation.

- In a Geiger counter, radiation produces charged particles in the gas contained in a tube, which generates an electrical current.
- The curie (Ci) and the becquerel (Bq) measure the activity, which is the number of nuclear transformations per second.

16.4 Half-Life of a Radioisotope

LEARNING GOAL: Given the half-life of a radioisotope, calculate the amount of radioisotope remaining after one or more half-lives.

- Every radioisotope has its own rate of emitting radiation.
- The time it takes for one-half of a radioactive sample to decay is called its half-life.
- For many medical radioisotopes, such as Tc-99m and I-131, half-lives are short.
- For other isotopes, usually naturally occurring ones such as C-14, Ra-226, and U-238, half-lives are extremely long.

16.5 Medical Applications Using Radioactivity

LEARNING GOAL: Describe the use of radioisotopes in medicine.

- In nuclear medicine, radioisotopes that go to specific sites in the body are given to the patient.
- By detecting the radiation they emit, an evaluation can be made about the location and extent of an injury, disease, tumor, or the level of function of a particular organ.
- Higher levels of radiation are used to treat or destroy tumors.
### 16.6 Nuclear Fission and Fusion

**LEARNING GOAL**
Describe the processes of nuclear fission and fusion.

- In fission, the bombardment of a large nucleus breaks it apart into smaller nuclei, releasing one or more types of radiation and a great amount of energy.
- In fusion, small nuclei combine to form larger nuclei while great amounts of energy are released.

### KEY TERMS

- **alpha particle** A nuclear particle identical to a helium nucleus, symbol $^0\text{He}$ or $^4\text{He}$.
- **becquerel (Bq)** A unit of activity of a radioactive sample equal to one disintegration per second.
- **beta particle** A particle identical to an electron, symbol $^0\text{e}$ or $^\beta$, that forms in the nucleus when a neutron changes to a proton and an electron.
- **chain reaction** A fission reaction that will continue once it has been initiated by a high-energy neutron bombarding a heavy nucleus such as uranium-235.
- **curie (Ci)** A unit of activity of a radioactive sample equal to $3.7 \times 10^{10}$ disintegrations/s.
- **decay curve** A diagram of the decay of a radioactive element.
- **equivalent dose** The measure of biological damage from an absorbed dose that has been adjusted for the type of radiation.
- **fission** A process in which large nuclei are split into smaller pieces, releasing large amounts of energy.
- **fusion** A reaction in which large amounts of energy are released when small nuclei combine to form larger nuclei.

### CORE CHEMISTRY SKILLS

*The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.*

### Writing Nuclear Equations (16.2)

- A nuclear equation is written with the atomic symbols of the original radioactive nucleus on the left, an arrow, and the new nucleus and the type of radiation emitted on the right.
- The sum of the mass numbers and the sum of the atomic numbers on one side of the arrow must equal the sum of the mass numbers and the sum of the atomic numbers on the other side.
- When an alpha particle is emitted, the mass number of the new nucleus decreases by 4, and its atomic number decreases by 2.
- When a beta particle is emitted, there is no change in the mass number of the new nucleus, but its atomic number increases by one.
- When a positron is emitted, there is no change in the mass number of the new nucleus, but its atomic number decreases by one.
- In gamma emission, there is no change in the mass number or the atomic number of the new nucleus.

**Example:**

**a.** Write a balanced nuclear equation for the alpha decay of Po-210.

**b.** Write a balanced nuclear equation for the beta decay of Co-60.

**Answer:**

**a.** When an alpha particle is emitted, we calculate the decrease of 4 in the mass number (210) of the polonium, and a decrease of 2 in its atomic number.

\[
\begin{align*}
\text{Po-210} & \rightarrow \text{Pb-206} + ^4\text{He} \\
\end{align*}
\]

Because lead has atomic number 82, the new nucleus must be an isotope of lead.

\[
\begin{align*}
\text{Po-210} & \rightarrow \text{Pb-206} + ^4\text{He} \\
\end{align*}
\]

**b.** When a beta particle is emitted, there is no change in the mass number (60) of the cobalt, but there is an increase of 1 in its atomic number.

\[
\begin{align*}
\text{Co-60} & \rightarrow \text{Ni-60} + ^0\text{e} \\
\end{align*}
\]

Because nickel has atomic number 28, the new nucleus must be an isotope of nickel.

\[
\begin{align*}
\text{Co-60} & \rightarrow \text{Ni-60} + ^0\text{e} \\
\end{align*}
\]
Using Half-Lives (16.4)

- The half-life of a radioisotope is the amount of time it takes for one-half of a sample to decay.
- The remaining amount of a radioisotope is calculated by dividing its quantity or activity by one-half for each half-life that has elapsed.

**Example:** Co-60 has a half-life of 5.3 yr. If the initial sample of Co-60 has an activity of 1200 Ci, what is its activity after 15.9 yr?

**Answer:**

In 15.9 yr, three half-lives have passed. Thus, the activity was reduced from 1200 Ci to 150 Ci.

<table>
<thead>
<tr>
<th>Years</th>
<th>Number of Half-Lives</th>
<th>Co-60 (Activity)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0</td>
<td>1200 Ci</td>
</tr>
<tr>
<td>5.3</td>
<td>1</td>
<td>600 Ci</td>
</tr>
<tr>
<td>10.6</td>
<td>2</td>
<td>300 Ci</td>
</tr>
<tr>
<td>15.9</td>
<td>3</td>
<td>150 Ci</td>
</tr>
</tbody>
</table>

In 15.9 yr, three half-lives have passed. Thus, the activity was reduced from 1200 Ci to 150 Ci.

---

**UNDERSTANDING THE CONCEPTS**

*The chapter sections to review are shown in parentheses at the end of each question.*

In problems 16.49 to 16.52, a nucleus is shown with protons and neutrons.

- **proton**
- **neutron**

16.49 Draw the new nucleus when this isotope emits a positron to complete the following figure: (16.2)

16.50 Draw the nucleus that emits a positron particle to complete the following figure: (16.2)

16.51 Draw the nucleus of the isotope that is bombarded in the following figure: (16.2)

16.52 Complete the following bombardment reaction by drawing the nucleus of the new isotope that is produced in the following figure: (16.2)

**16.53** Carbon dating of small bits of plant colouring used in cave paintings is used to determine the age of the paintings. Experts compared the ratio of Carbon-12 and Carbon-14 in the painting to estimate its age. It is found that the painting has 6.25% Carbon-14 in comparison to the sample of the living plant. Given that Carbon-14 has a half-life of 5730 year, what is the age of the painting? (16.4)

**16.54** Use the following decay curve for iodine-131 to answer questions a to c: (16.4)

a. Complete the values for the mass of radioactive iodine-131 on the vertical axis.

b. Complete the number of days on the horizontal axis.

c. What is the half-life, in days, of iodine-131?
16.55 Determine the number of protons and number of neutrons in the nucleus of each of the following: (16.1)
   a. magnesium-26
   b. strontium-88
   c. iodine-131
   d. copper-64

16.56 Determine the number of protons and number of neutrons in the nucleus of each of the following: (16.1)
   a. neon-22
   b. cobalt-60
   c. tin-122
   d. manganese-56

16.57 Identify each of the following as alpha decay, beta decay, positron emission, or gamma emission: (16.1, 16.2)
   a. $^{223}_{88}$Ra $\rightarrow^{221}_{86}$Rn + $^{2}_{4}$He
   b. $^{60}_{27}$Co $\rightarrow^{60}_{26}$Ni + $^{0}_{-1}$e
   c. $^{18}_{9}$F $\rightarrow^{18}_{8}$O + $^{0}_{+1}$e

16.58 Identify each of the following as alpha decay, beta decay, positron emission, or gamma emission: (16.1, 16.2)
   a. $^{249}_{92}$Pu $\rightarrow^{247}_{90}$U + $^{2}_{4}$He
   b. $^{131}_{53}$I $\rightarrow^{131}_{52}$I + $^{0}_{0}$e
   c. $^{74}_{31}$Na $\rightarrow^{74}_{30}$Ne + $^{0}_{+1}$e

16.59 Write the balanced nuclear equation for each of the following: (16.1, 16.2)
   a. Th-225 ($\alpha$ decay)
   b. Bi-210 ($\alpha$ decay)
   c. cesium-137 ($\beta$ decay)
   d. tin-126 ($\beta$ decay)
   e. F-18 ($\beta^+$ emission)

16.60 Write the balanced nuclear equation for each of the following: (16.1, 16.2)
   a. potassium-40 ($\beta$ decay)
   b. sulfur-35 ($\beta$ decay)
   c. platinum-190 ($\alpha$ decay)
   d. Ra-210 ($\alpha$ decay)
   e. In-113m ($\gamma$ emission)

16.61 Complete each of the following nuclear equations: (16.2)
   a. $^{1}_{0}$n + $^{19}_{9}$F $\rightarrow^{20}_{9}$F + $^{0}_{+1}$e
   b. $^{3}_{2}$He + $^{39}_{19}$K $\rightarrow^{40}_{20}$Ca + $^{0}_{0}$e
   c. $^{1}_{0}$n + ? $\rightarrow^{23}_{12}$Mg + $^{4}_{2}$He
   d. $^{13}_{53}$Cs $\rightarrow^{13}_{52}$Ba + $^{0}_{0}$e

16.62 Complete each of the following nuclear equations: (16.2)
   a. $^{1}_{0}$n + $^{12}_{6}$B $\rightarrow^{13}_{6}$B
   b. $^{2}_{2}$He + $^{124}_{50}$Sn $\rightarrow^{124}_{50}$Sn + $^{0}_{0}$e
   c. $^{3}_{1}$e + ? $\rightarrow^{5}_{2}$Li + $^{3}_{0}$He
   d. $^{5}_{2}$He + $^{232}_{90}$Th $\rightarrow^{229}_{88}$Ac + 2 $^{0}_{0}$n

16.63 Write the balanced nuclear equation for each of the following: (16.2)
   a. When two oxygen-16 atoms collide, one of the products is an alpha particle.
   b. When californium-249 is bombarded by oxygen-18, a new element, seaborgium-263, and four neutrons are produced.
   c. Radon-222 undergoes alpha decay.
   d. An atom of strontium-80 emits a positron.

16.64 Write the balanced nuclear equation for each of the following: (16.2)
   a. Polonium-210 decays to lead-206.
   b. Bismuth-211 emits an alpha particle.
   c. A radioisotope emits a positron to form titanium-48.
   d. An atom of germanium-69 emits a positron.

16.65 A 120-mg sample of technetium-99m is used for a diagnostic test. If technetium-99m has a half-life of 6.0 h, how many milligrams of the technetium-99m sample remains active 24 h after the test? (16.4)

16.66 The half-life of oxygen-15 is 124 s. If a sample of oxygen-15 has an activity of 4000 Bq, how many minutes will elapse before it has an activity of 500 Bq? (16.4)

16.67 What is the difference between fission and fusion? (16.6)

16.68 a. What are the products in the fission of uranium-235 that make possible a nuclear chain reaction? (16.6)
   b. What is the purpose of placing control rods among uranium samples in a nuclear reactor? (16.6)

16.69 Where does fusion occur naturally? (16.6)

16.70 Why are scientists continuing to try to build a fusion reactor even though the very high temperatures it requires have been difficult to reach and maintain? (16.6)

Applications

16.71 The activity of K-40 in a 50.-kg human body is estimated to be 110 nCi. What is the activity in becquerels? (16.3)

16.72 The activity of C-14 in a 70.-kg human body is estimated to be 3.7 kBq. What is this activity in microcuries? (16.3)

16.73 If the amount of radioactive phosphorus-32, used to treat leukemia, in a sample decreases from 1.2 mg to 0.30 mg in 28.6 days, what is the half-life of phosphorus-32? (16.4)

16.74 If the amount of radioactive iodine-123, used to treat thyroid cancer, in a sample decreases from 0.4 mg to 0.1 mg in 26.4 h, what is the half-life of iodine-123? (16.4)

16.75 Calcium-47, used to evaluate bone metabolism, has a half-life of 4.5 days. (16.2, 16.4)
   a. Write the balanced nuclear equation for the beta decay of calcium-47.
   b. How many milligrams of a 16-mg sample of calcium-47 remain after 18 days?
   c. How many days have passed if 4.8 mg of calcium-47 decayed to 1.2 mg of calcium-47?

16.76 Cesium-137, used in cancer treatment, has a half-life of 30 yr. (16.2, 16.4)
   a. Write the balanced nuclear equation for the beta decay of cesium-137.
   b. How many milligrams of a 16-mg sample of cesium-137 remain after 90 yr?
   c. How many years are required for 28 mg of cesium-137 to decay to 3.5 mg of cesium-137?
CHALLENGE QUESTIONS

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

16.77 Write the balanced nuclear equation for each of the following radioactive emissions: (16.2)
   a. an alpha particle from Ag-111
   b. a beta particle from Mn-55
   c. a positron from Ca-43

16.78 Write the balanced nuclear equation for each of the following radioactive emissions: (16.2)
   a. an alpha particle from Po-208
   b. a beta particle from I-131
   c. a positron from Al-25

16.79 All the elements beyond uranium, the transuranium elements, have been prepared by bombardment and are not naturally occurring elements. The first transuranium element neptunium, Np, was prepared by bombarding U-238 with neutrons to form a neptunium atom and a beta particle. Complete the following equation: (16.2)
   \[ _{92}^{238}U + \_n \rightarrow \_Np + \_\beta \]

16.80 One of the most recent transuranium elements ununoctium-294 (Uuo-294), atomic number 118, was prepared by bombarding californium-249 with another isotope. Complete the following equation for the preparation of this new element: (16.2)
   \[ _{98}^{249}Cf + \_n \rightarrow \_\text{Uuo} + \_\beta \]

16.81 A 64-\( \mu \text{Ci} \) sample of Tl-201 decays to 4.0 \( \mu \text{Ci} \) in 12 days. What is the half-life, in days, of Tl-201? (16.3, 16.4)

16.82 A wooden object from the site of an ancient temple has a carbon-14 activity of 10 counts/min compared with a reference piece of wood cut today that has an activity of 40 counts/min. If the half-life for carbon-14 is 5730 yr, what is the age of the ancient wood object? (16.3, 16.4)

16.83 The half-life for the radioactive decay of calcium-47 is 4.5 days. If a sample has an activity of 1.0 \( \mu \text{Ci} \) after 27 days, what was the initial activity, in microcuries, of the sample? (16.3, 16.4)

16.84 The half-life for the radioactive decay of Ce-141 is 32.5 days. If a sample has an activity of 4.0 \( \mu \text{Ci} \) after 130 days have elapsed, what was the initial activity, in microcuries, of the sample? (16.3, 16.4)

16.85 Scientists created a new synthetic element californium, Cf-246, by bombarding U-238 with C-14. Other than Cf-246, 4 neutrons are produced in the reaction. Write the balanced nuclear equation for the synthesis of californium. (16.2)

16.86 Nuclear fission of U-235 can be induced by the bombardment of neutron to give Ba-141, Kr-92, and three neutrons. Write the balanced nuclear equation for the reaction and explain why a chain reaction will occur after the nuclear fission of one U-235. (16.2)

Applications

16.87 A nuclear technician was accidentally exposed to potassium-42 while doing brain scans for possible tumors. The error was not discovered until 36 h later when the activity of the potassium-42 sample was 2.0 \( \mu \text{Ci} \). If potassium-42 has a half-life of 12 h, what was the activity of the sample at the time the technician was exposed? (16.3, 16.4)

16.88 The radioisotope sodium-24 is used to determine the levels of electrolytes in the body. A 16-\( \mu \text{g} \) sample of sodium-24 decays to 2.0 \( \mu \text{g} \) in 45 h. What is the half-life, in hours, of sodium-24? (16.4)

ANSWERS

Answers to Selected Questions and Problems

16.1 a. alpha particle
   b. positron
   c. gamma radiation

16.3 a. \( ^{39}_{19}\text{K}, \quad ^{40}_{20}\text{K}, \quad ^{41}_{20}\text{K} \)
   b. They all have 19 protons and 19 electrons, but they differ in the number of neutrons.

16.5 a. \( \beta \) or \( ^{0}_{-1}\text{e} \)
   b. \( \alpha \) or \( ^{4}_{2}\text{He} \)
   c. \( n \) or \( ^{0}_{1}\text{p} \)
   d. \( ^{38}_{18}\text{Ar} \)
   e. \( ^{14}_{6}\text{C} \)

16.7 a. \( ^{64}_{32}\text{Cu} \)
   b. \( ^{75}_{34}\text{Se} \)
   c. \( ^{25}_{11}\text{Na} \)
   d. \( ^{15}_{7}\text{N} \)

16.9

<table>
<thead>
<tr>
<th>Medical Use</th>
<th>Atomic Symbol</th>
<th>Mass Number</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Heart imaging</td>
<td>( ^{201}_{81}\text{Tl} )</td>
<td>201</td>
<td>81</td>
<td>120</td>
</tr>
<tr>
<td>Radiation therapy</td>
<td>( ^{60}_{27}\text{Co} )</td>
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<td>27</td>
<td>33</td>
</tr>
<tr>
<td>Abdominal scan</td>
<td>( ^{67}_{31}\text{Ga} )</td>
<td>67</td>
<td>31</td>
<td>36</td>
</tr>
<tr>
<td>Hyperthyroidism</td>
<td>( ^{131}_{53}\text{I} )</td>
<td>131</td>
<td>53</td>
<td>78</td>
</tr>
<tr>
<td>Leukemia treatment</td>
<td>( ^{32}_{15}\text{P} )</td>
<td>32</td>
<td>15</td>
<td>17</td>
</tr>
</tbody>
</table>

16.11 a. 1, alpha particle
   b. 3, gamma radiation

16.12 a. 1, alpha particle
   c. 1, alpha particle
16.13  a. $^{208}_{84}\text{Po} \rightarrow ^{204}_{82}\text{Pb} + ^{4}_{2}\text{He}$
b. $^{222}_{89}\text{Th} \rightarrow ^{222}_{88}\text{Ra} + ^{4}_{2}\text{He}$
c. $^{214}_{92}\text{Fr} \rightarrow ^{214}_{91}\text{Fr} + ^{4}_{2}\text{He}$
d. $^{220}_{86}\text{Rn} \rightarrow ^{216}_{84}\text{Po} + ^{4}_{2}\text{He}$
16.15  a. $^{26}_{11}\text{Na} \rightarrow ^{27}_{12}\text{Mg} + ^{0}_{-1}\text{e}$
b. $^{16}_{8}\text{O} \rightarrow ^{16}_{8}\text{F} + ^{0}_{0}\text{e}$
c. $^{92}_{40}\text{Sr} \rightarrow ^{82}_{38}\text{Y} + ^{0}_{-2}\text{e}$
d. $^{56}_{26}\text{Fe} \rightarrow ^{56}_{26}\text{Fe} + ^{0}_{0}\text{e}$
16.17  a. $^{28}_{14}\text{Si} \rightarrow ^{28}_{13}\text{Al} + ^{0}_{0}\text{e}$
b. $^{54}_{27}\text{Co} \rightarrow ^{54}_{26}\text{Fe} + ^{0}_{-2}\text{e}$
c. $^{87}_{37}\text{Rb} \rightarrow ^{87}_{36}\text{Sr} + ^{0}_{-1}\text{e}$
d. $^{82}_{31}\text{Rh} \rightarrow ^{82}_{32}\text{Rn} + ^{0}_{+1}\text{e}$
16.19  a. $^{14}_{6}\text{Si}$, beta decay
b. $^{19}_{0}\text{Y}$, gamma emission
c. $^{0}_{-1}\text{e}$, beta decay
d. $^{238}_{92}\text{U}$, alpha decay
e. $^{189}_{79}\text{Au}$, positron emission
16.21  a. $^{13}_{7}\text{Be}$
b. $^{15}_{7}\text{Cl}$
c. $^{17}_{7}\text{Al}$
d. $^{17}_{8}\text{O}$
16.23  a. 2, absorbed dose
b. 3, biological damage
c. 1, activity
d. 2, absorbed dose
16.25  The technician exposed to 5 rad received the higher amount of radiation.
16.27  a. 294 $\mu$Ci
b. 0.5 Gy
16.29  a. two half-lives
b. one half-life
c. three half-lives
16.31  a. 40.0 mg
b. 20.0 mg
c. 10.0 mg
d. 5.00 mg
16.33  a. 130 days
b. 195 days
16.35  a. Because the elements Ca and P are part of the bone, the radioactive isotopes of Ca and P will become part of the bony structures of the body, where their radiation can be used to diagnose or treat bone diseases.
b. Strontium (Sr) acts much like calcium (Ca) because both are Group 2A (2) elements. The body will accumulate radioactive strontium in bones in the same way that it incorporates calcium. Radioactive strontium is harmful to children because the radiation it produces causes more damage in cells that are dividing rapidly.
16.37  180 $\mu$Ci
16.39  Nuclear fission is the splitting of a large atom into smaller fragments with the release of large amounts of energy.
16.41  $^{103}_{42}\text{Mo}$
16.43  a. fission
b. fusion
c. fission
d. fusion
16.45  a. $2.0 \times 10^{-3}$ Ci
b. 2.0 mCi
16.47  9.0 days
16.51  
16.53  22,920 yr old
16.55  a. 12 protons and 14 neutrons
b. 38 protons and 50 neutrons
c. 53 protons and 78 neutrons
d. 29 protons and 35 neutrons
16.57  a. alpha decay
b. beta decay
c. positron emission
16.59  a. $^{222}_{87}\text{Th} \rightarrow ^{222}_{88}\text{Ra} + ^{4}_{2}\text{He}$
b. $^{210}_{82}\text{Bi} \rightarrow ^{206}_{81}\text{Tl} + ^{4}_{2}\text{He}$
c. $^{137}_{55}\text{Cs} \rightarrow ^{137}_{56}\text{Ba} + ^{0}_{0}\text{e}$
d. $^{128}_{50}\text{Sn} \rightarrow ^{128}_{51}\text{Sb} + ^{0}_{-1}\text{e}$
e. $^{18}_{0}\text{F} \rightarrow ^{18}_{8}\text{O} + ^{0}_{+1}\text{e}$
16.61  a. $^{20}_{8}\text{O}$
b. $^{41}_{21}\text{Ca}$
c. $^{20}_{9}\text{F}$
d. $^{13}_{5}\text{Cl}$
16.63  a. $^{15}_{8}\text{O} + ^{15}_{9}\text{O} \rightarrow ^{30}_{16}\text{Si} + ^{0}_{0}\text{e}$
b. $^{18}_{8}\text{O} + ^{24}_{12}\text{Cf} \rightarrow ^{26}_{16}\text{Sg} + ^{4}_{0}\text{He}$
c. $^{222}_{89}\text{Rn} \rightarrow ^{218}_{86}\text{Po} + ^{3}_{2}\text{He}$
d. $^{80}_{36}\text{Sr} \rightarrow ^{80}_{37}\text{Rb} + ^{0}_{+1}\text{e}$
16.65  7.5 mg of Tc-99m
16.67  In the fission process, an atom splits into smaller nuclei. In fusion, small nuclei combine (fuse) to form a larger nucleus.
16.69  Fusion occurs naturally in the Sun and other stars.
16.71  $4.07 \times 10^{3}$ Bq
16.73  14.3 days
16.75  a. $^{47}_{22}\text{Ca} \rightarrow ^{47}_{23}\text{Sc} + ^{0}_{-1}\text{e}$
b. 1.0 mg of Ca-47
c. 9.0 days
16.77  a. $^{111}_{47}\text{Ag} \rightarrow ^{107}_{45}\text{Rh} + ^{4}_{2}\text{He}$
b. $^{55}_{25}\text{Mn} \rightarrow ^{55}_{26}\text{Fe} + ^{0}_{-2}\text{e}$
c. $^{43}_{22}\text{Ca} \rightarrow ^{43}_{23}\text{K} + ^{0}_{+1}\text{e}$
16.79  $^{14}_{0}\text{He} + ^{238}_{92}\text{U} \rightarrow ^{238}_{93}\text{Np} + ^{0}_{-1}\text{e}$
16.81  3.0 days
16.83  64 $\mu$Ci
16.85  $^{12}_{6}\text{C} + ^{238}_{92}\text{U} \rightarrow ^{244}_{98}\text{Cf} + ^{4}_{0}\text{He}$
16.87  16 $\mu$Ci
**CL.33** Consider the reaction of sodium oxalate (Na$_2$C$_2$O$_4$) and potassium permanganate (KMnO$_4$) in acidic solution. The unbalanced equation is the following: (9.2, 9.3, 12.4, 12.5, 15.2)

\[ \text{MnO}_4^- (aq) + C_2O_4^{2-} (aq) \rightarrow \text{Mn}^{2+} (aq) + CO_2(g) \]

In an oxidation-reduction titration, the KMnO$_4$ from the buret reacts with Na$_2$C$_2$O$_4$.

a. What is the balanced oxidation half-reaction?

b. What is the balanced reduction half-reaction?

c. What is the balanced ionic equation for the reaction?

d. If 24.6 mL of a KMnO$_4$ solution is needed to titrate a solution containing 0.758 g of sodium oxalate (Na$_2$C$_2$O$_4$), what is the molarity of the KMnO$_4$ solution?

**CL.34** A strip of magnesium metal dissolves rapidly in 6.00 mL of a 0.150 M HCl solution, producing hydrogen gas and magnesium chloride. (9.2, 9.3, 12.4, 12.6, 15.1, 15.2)

\[ \text{Mg}(s) + \text{HCl}(aq) \rightarrow \text{H}_2(g) + \text{MgCl}_2(aq) \] Unbalanced

Magnesium metal reacts vigorously with hydrochloric acid.

a. Assign oxidation numbers to all of the elements in the reactants and products.

b. What is the balanced chemical equation for the reaction?

c. What is the oxidizing agent?

d. What is the reducing agent?

e. What is the pH of the 0.150 M HCl solution?

f. How many grams of magnesium can dissolve in the HCl solution?

**CL.35** A piece of magnesium with a mass of 0.121 g is added to 50.0 mL of a 1.00 M HCl solution at a temperature of 22 °C. When the magnesium dissolves, the solution reaches a temperature of 33 °C. For the equation, see your answer to problem CL.34b. (9.2, 9.3, 9.4, 9.6, 12.4, 12.6)

a. What is the limiting reactant?

b. What volume, in milliliters, of hydrogen gas would be produced at 33 °C when the pressure is 750. mmHg?

c. How many joules were released by the reaction of the magnesium? Assume the density of the HCl solution is 1.00 g/mL and the specific heat of the HCl solution is the same as that of water.

d. What is the heat of reaction for magnesium in joules/gram? in kilojoules/mol?

**CL.36** The iceman known as Ötzi was discovered in a high mountain pass on the Austrian–Italian border. Samples of his hair and bones had carbon-14 activity that was 50% of that present in new hair or bone. Carbon-14 undergoes beta decay and has a half-life of 5730 yr. (16.2, 16.4)

a. How long ago did Ötzi live?

b. Write a balanced nuclear equation for the decay of carbon-14.

**CL.37** Some of the isotopes of silicon are listed in the following table: (4.4, 5.4, 10.1, 10.3, 16.2, 16.4)

<table>
<thead>
<tr>
<th>Isotope</th>
<th>% Natural Abundance</th>
<th>Atomic Mass</th>
<th>Half-Life (radioactive)</th>
<th>Radiation</th>
</tr>
</thead>
<tbody>
<tr>
<td>27Si</td>
<td>26.987</td>
<td>4.9 s</td>
<td>Positron</td>
<td></td>
</tr>
<tr>
<td>28Si</td>
<td>92.230</td>
<td>27.977</td>
<td>Stable</td>
<td></td>
</tr>
<tr>
<td>29Si</td>
<td>4.683</td>
<td>28.976</td>
<td>Stable</td>
<td></td>
</tr>
<tr>
<td>30Si</td>
<td>3.087</td>
<td>29.974</td>
<td>Stable</td>
<td></td>
</tr>
<tr>
<td>31Si</td>
<td>30.975</td>
<td>2.6 h</td>
<td>Beta</td>
<td></td>
</tr>
</tbody>
</table>

a. In the following table, indicate the number of protons, neutrons, and electrons for each isotope listed:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>17Si</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>28Si</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>29Si</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>30Si</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>31Si</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

b. What are the electron configuration and the abbreviated electron configuration of silicon?

c. Calculate the atomic mass for silicon using the isotopes that have a natural abundance.

d. Write the balanced nuclear equations for the positron emission of Si-27 and the beta decay of Si-31.

e. Draw the Lewis structure and predict the shape of SiCl$_4$.

f. How many hours are needed for a sample of Si-31 with an activity of 16 μCi to decay to 2.0 μCi?
CL.38  K\(^+\) is an electrolyte required by the human body and found in many foods as well as salt substitutes. One of the isotopes of potassium is potassium-40, which has a natural abundance of 0.012% and a half-life of 1.30 \(\times\) 10\(^{9}\) yr. The isotope potassium-40 decays to calcium-40 or to argon-40. A typical activity for potassium-40 is 7.0 \(\mu\)Ci per gram. (16.2, 16.3, 16.4)

![Potassium chloride is used as a salt substitute.](image)

a. Write a balanced nuclear equation for each type of decay.

b. Identify the particle emitted for each type of decay.

c. How many K\(^+\) ions are in 3.5 oz of KCl?

d. What is the activity of 25 g of KCl in becquerels?

CL.39  Uranium-238 decays in a series of nuclear changes until stable lead-206 is produced. Complete the following nuclear equations that are part of the uranium-238 decay series: (16.2, 16.3, 16.4)

a. \(^{238}\text{U} \rightarrow ^{234}\text{Th} + ?

b. \(^{234}\text{Th} \rightarrow ? + ^{0}\text{He}

c. ? \rightarrow ^{222}\text{Rn} + ^{4}\text{He}

CL.40  Of much concern to environmentalists is radon-222, which is a radioactive noble gas that can seep from the ground into basements of homes and buildings. Radon-222 is a product of the decay of radium-226 that occurs naturally in rocks and soil in much of the United States. Radon-222, which has a half-life of 3.8 days, decays by emitting an alpha particle. Radon-222, which is a gas, can be inhaled into the lungs where it is strongly associated with lung cancer. Radon levels in a home can be measured with a home radon-detection kit. Environmental agencies have set the maximum level of radon-222 in a home at 4 picocuries per liter (pCi/L) of air. (11.7, 16.2, 16.3, 16.4)

![A home detection kit is used to measure the level of radon-222.](image)

a. Write the balanced nuclear equation for the decay of Ra-226.

b. Write the balanced nuclear equation for the decay of Rn-222.

c. If a room contains 24 000 atoms of radon-222, how many atoms of radon-222 remain after 15.2 days?

d. Suppose a room has a volume of 72 000 L (7.2 \(\times\) 10\(^{4}\) L). If the radon level is the maximum allowed (4 pCi/L), how many alpha particles are emitted from Rn-222 in one day? (1 Ci = 3.7 \(\times\) 10\(^{10}\) disintegrations per second)

ANSWERS

CL.33  a. \(\text{C}_2\text{O}_4^{2-}(aq) \rightarrow 2\text{CO}_2(g) + 2\ e^-\)

b. \(5\ e^- + 8\text{H}^+(aq) + \text{MnO}_4^{-}(aq) \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l)\)

c. \(16\text{H}^+(aq) + 2\text{MnO}_4^{-}(aq) + 5\text{C}_2\text{O}_4^{2-}(aq) \rightarrow 10\text{CO}_2(g) + 2\text{Mn}^{2+}(aq) + 8\text{H}_2\text{O}(l)\)

d. 0.0920 M KMnO\(_4\) solution

CL.35  a. Mg is the limiting reactant.

b. 127 mL of \(\text{H}_2(g)\)

c. 2.3 \(\times\) 10\(^{9}\) J

d. 1.90 \(\times\) 10\(^4\) J/g; 462 kJ/mol

CL.37  a. | Isotope | Number of Protons | Number of Neutrons | Number of Electrons |
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>(^{28}\text{Si})</td>
<td>14</td>
<td>13</td>
<td>14</td>
</tr>
<tr>
<td>(^{30}\text{Si})</td>
<td>14</td>
<td>14</td>
<td>14</td>
</tr>
<tr>
<td>(^{32}\text{Si})</td>
<td>14</td>
<td>15</td>
<td>14</td>
</tr>
<tr>
<td>(^{34}\text{Si})</td>
<td>14</td>
<td>16</td>
<td>14</td>
</tr>
<tr>
<td>(^{36}\text{Si})</td>
<td>14</td>
<td>17</td>
<td>14</td>
</tr>
</tbody>
</table>

b. \(1s^22s^22p^63s^23p^2; [\text{Ne}]3s^23p^2\)

c. 28.09 amu

d. \(^{28}\text{Si} \rightarrow ^{27}\text{Al} + ^{0}\text{He}\)

e. \(^{30}\text{Si} \rightarrow ^{30}\text{P} + ^{0}\text{He}\)

f. 7.8 h
CarEEr

Firefighter/Emergency medical technician

Firefighters/emergency medical technicians are first responders to fires, accidents, and other emergency situations. They are required to have an emergency medical technician certification in order to be able to treat seriously injured people. By combining the skills of a firefighter and an emergency medical technician, they increase the survival rates of the injured. The physical demands of firefighters are extremely high as they fight, extinguish, and prevent fires while wearing heavy protective clothing. They also train for and participate in firefighting drills, and maintain fire equipment so that it is always working and ready. Firefighters must also be knowledgeable about fire codes, arson, and the handling and disposal of hazardous materials. Since firefighters also provide emergency care for sick and injured people, they need to be aware of emergency medical and rescue procedures, as well as the proper methods for controlling the spread of infectious disease.

AT 4:35 A.M., A RESCUE crew responded to a call about a house fire. At the scene, Jack, a firefighter/emergency medical technician (EMT) found Diane, a 62-year old woman, lying in the front yard of her house. In his assessment, Jack reported that Diane had second- and third-degree burns over 40% of her body as well as a broken leg. He placed an oxygen re-breather mask on Diane to provide a high concentration of oxygen. Another firefighter/EMT, Nancy, began dressing the burns with sterile water and cling film, a first aid material made of polyvinyl chloride, which does not stick to the skin and is protective. Jack and his crew transported Diane to the burn center for further treatment.

At the scene of the fire, arson investigators used trained dogs to find traces of accelerants and fuel. Gasoline, which is often found at arson scenes, is a mixture of organic molecules called alkanes. Alkanes or hydrocarbons are chains of carbon and hydrogen atoms. The alkanes present in gasoline consist of a mixture of 5 to 8 carbon atoms in a chain. Alkanes are extremely combustible; they react with oxygen to form carbon dioxide, water, and large amounts of heat. Because alkanes undergo combustion reactions, they can be used to start arson fires.

Organic Chemistry

CAREER

Firefighter/Emergency Medical Technician

Firefighters/emergency medical technicians are first responders to fires, accidents, and other emergency situations. They are required to have an emergency medical technician certification in order to be able to treat seriously injured people. By combining the skills of a firefighter and an emergency medical technician, they increase the survival rates of the injured. The physical demands of firefighters are extremely high as they fight, extinguish, and prevent fires while wearing heavy protective clothing. They also train for and participate in firefighting drills, and maintain fire equipment so that it is always working and ready. Firefighters must also be knowledgeable about fire codes, arson, and the handling and disposal of hazardous materials. Since firefighters also provide emergency care for sick and injured people, they need to be aware of emergency medical and rescue procedures, as well as the proper methods for controlling the spread of infectious disease.
17.1 Alkanes

**LEARNING GOAL** Write the IUPAC names and draw the condensed or line-angle structural formulas for alkanes.

At the beginning of the nineteenth century, scientists classified chemical compounds as inorganic or organic. An inorganic compound was a substance that was composed of minerals, and an organic compound was a substance that came from an organism, thus the use of the word organic. It was thought that some type of “vital force,” which could only be found in living cells, was required to synthesize an organic compound. This perception was shown to be incorrect in 1828, when the German chemist Friedrich Wöhler synthesized urea, a product of protein metabolism, by heating an inorganic compound, ammonium cyanate. Organic compounds contain carbon and hydrogen and sometimes oxygen and other nonmetallic elements.

![Ammonium cyanate](image1.png)

We use many organic compounds every day in products such as fuels, clothing, drugs, and cosmetics. The foods in our diets are composed of organic compounds such as carbohydrates, fats, and proteins that undergo digestion and metabolism to give us energy and small organic compounds that are used to build and repair the cells of our bodies.

The formulas of organic compounds are written with carbon first, followed by hydrogen, and then any other elements. Organic compounds typically have low melting and boiling points, are not soluble in water, and are less dense than water. For example, vegetable oil, which is a mixture of organic compounds, does not dissolve in water but floats on top. Many organic compounds undergo combustion and burn vigorously in air. By contrast, many inorganic compounds have high melting and boiling points. Inorganic compounds that are ionic are usually soluble in water, and most do not burn in air. **TABLE 17.1** contrasts some of the properties associated with organic and inorganic compounds, such as propane, C₃H₈, and sodium chloride, NaCl (see **FIGURE 17.1**).
Vegetable oil, a mixture of organic compounds, is not soluble in water.

**TABLE 17.1 Some Properties of Organic and Inorganic Compounds**

<table>
<thead>
<tr>
<th>Property</th>
<th>Organic</th>
<th>Example: $\text{C}_3\text{H}_8$</th>
<th>Inorganic</th>
<th>Example: $\text{NaCl}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Elements Present</td>
<td>C and H, sometimes O, S, N, P, or Cl (F, Br, I)</td>
<td>C and H</td>
<td>Most metals and nonmetals</td>
<td>Na and Cl</td>
</tr>
<tr>
<td>Particles</td>
<td>Molecules</td>
<td>$\text{C}_3\text{H}_8$</td>
<td>Mostly ions</td>
<td>$\text{Na}^+$ and $\text{Cl}^-$</td>
</tr>
<tr>
<td>Bonding</td>
<td>Mostly covalent</td>
<td>Covalent</td>
<td>Many are ionic, some covalent</td>
<td>Ionic</td>
</tr>
<tr>
<td>Polarity of Bonds</td>
<td>Nonpolar, unless a strongly electronegative atom is present</td>
<td>Nonpolar</td>
<td>Most are ionic or polar covalent, a few are nonpolar covalent</td>
<td>Ionic</td>
</tr>
<tr>
<td>Melting Point</td>
<td>Usually low</td>
<td>$-188 , ^\circ \text{C}$</td>
<td>Usually high</td>
<td>$801 , ^\circ \text{C}$</td>
</tr>
<tr>
<td>Boiling Point</td>
<td>Usually low</td>
<td>$-42 , ^\circ \text{C}$</td>
<td>Usually high</td>
<td>$1413 , ^\circ \text{C}$</td>
</tr>
<tr>
<td>Flammability</td>
<td>High</td>
<td>Burns in air</td>
<td>Low</td>
<td>Does not burn</td>
</tr>
<tr>
<td>Solubility in Water</td>
<td>Not soluble unless a polar group is present</td>
<td>No</td>
<td>Most are soluble unless nonpolar</td>
<td>Yes</td>
</tr>
</tbody>
</table>

**FIGURE 17.1** Propane, $\text{C}_3\text{H}_8$, is an organic compound, whereas sodium chloride, $\text{NaCl}$, is an inorganic compound. Why is propane used as a fuel?

**Representations of Carbon Compounds**

**Hydrocarbons** are organic compounds that consist of only carbon and hydrogen. In organic molecules, every carbon atom has four bonds. In the simplest hydrocarbon, methane ($\text{CH}_4$), the carbon atom forms an octet by sharing its four valence electrons with four hydrogen atoms.

$$\cdot\text{C}\cdot + 4\cdot \rightarrow \text{H}\cdot\text{C}:\text{H} \quad \rightarrow \quad \text{H} = \text{H} - \text{C} - \text{H}$$

Methane
The most accurate representation of methane is the three-dimensional *space-filling model* (a) in which spheres show the relative size and shape of all the atoms. Another type of three-dimensional representation is the *ball-and-stick model* (b), where the atoms are shown as balls and the bonds between them are shown as sticks. In the ball-and-stick model of methane, CH₄, the covalent bonds from the carbon atom to each hydrogen atom are directed to the corners of a tetrahedron with bond angles of 109°. In the *wedge–dash model* (c), the three-dimensional shape is represented by symbols of the atoms with lines for bonds in the plane of the page, wedges for bonds that project out from the page, and dashes for bonds that are behind the page.

However, the three-dimensional models are awkward to draw and view for more complex molecules. Therefore, it is more practical to use their corresponding two-dimensional formulas. The **expanded structural formula** (d) shows all of the atoms and the bonds connected to each atom. A **condensed structural formula** (e) shows the carbon atoms each grouped with the attached number of hydrogen atoms.

The hydrocarbon ethane with two carbon atoms and six hydrogen atoms can be represented by a similar set of three- and two-dimensional models and formulas in which each carbon atom is bonded to another carbon and three hydrogen atoms. As in methane, each carbon atom in ethane retains a tetrahedral shape. A hydrocarbon is referred to as a **saturated hydrocarbon** when all the bonds in the molecule are single bonds.

### Naming Alkanes

More than 90% of the compounds in the world are organic compounds. The large number of carbon compounds is possible because the covalent bond between carbon atoms is very strong, allowing carbon atoms to form long, stable chains.

The **alkanes** are a type of hydrocarbon in which the carbon atoms are connected only by single bonds. One of the most common uses of alkanes is as fuels. Methane, used in gas heaters and gas cooktops, is an alkane with one carbon atom. The alkanes ethane, propane, and butane contain two, three, and four carbon atoms, respectively, connected in a row or a **continuous chain**. As we can see, the names for alkanes end in *ane*. Such names are part of the **IUPAC** (International Union of Pure and Applied Chemistry) system used by chemists to name organic compounds. Alkanes with five or more carbon atoms in a chain are named using Greek prefixes: *pent* (5), *hex* (6), *hept* (7), *oct* (8), *non* (9), and *dec* (10) (see **TABLE 17.2**).
TABLE 17.2 IUPAC Names, Molecular Formulas, Condensed and Line-Angle Structural Formulas of the First Ten Alkanes

<table>
<thead>
<tr>
<th>Carbon Atoms</th>
<th>IUPAC Name</th>
<th>Molecular Formula</th>
<th>Condensed Structural Formula</th>
<th>Line-Angle Structural Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Methane</td>
<td>CH₄</td>
<td>CH₄</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>Ethane</td>
<td>C₂H₆</td>
<td>CH₃—CH₃</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>Propane</td>
<td>C₃H₈</td>
<td>CH₃—CH₂—CH₃</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>Butane</td>
<td>C₄H₁₀</td>
<td>CH₃—CH₂—CH₃—CH₃</td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>Pentane</td>
<td>C₅H₁₂</td>
<td>CH₃—CH₂—CH₂—CH₂—CH₃</td>
<td></td>
</tr>
<tr>
<td>6</td>
<td>Hexane</td>
<td>C₆H₁₄</td>
<td>CH₃—CH₂—CH₂—CH₂—CH₂—CH₃</td>
<td></td>
</tr>
<tr>
<td>7</td>
<td>Heptane</td>
<td>C₇H₁₆</td>
<td>CH₃—CH₂—CH₂—CH₂—CH₂—CH₂—CH₃</td>
<td></td>
</tr>
<tr>
<td>8</td>
<td>Octane</td>
<td>C₈H₁₈</td>
<td>CH₃—CH₂—CH₂—CH₂—CH₂—CH₂—CH₂—CH₃</td>
<td></td>
</tr>
<tr>
<td>9</td>
<td>Nonane</td>
<td>C₉H₂₀</td>
<td>CH₃—CH₂—CH₂—CH₂—CH₂—CH₂—CH₂—CH₂—CH₃</td>
<td></td>
</tr>
<tr>
<td>10</td>
<td>Decane</td>
<td>C₁₀H₂₂</td>
<td>CH₃—CH₂—CH₂—CH₂—CH₂—CH₂—CH₂—CH₂—CH₂—CH₂—CH₃</td>
<td></td>
</tr>
</tbody>
</table>

Condensed and Line-Angle Structural Formulas

In a condensed structural formula, each carbon atom and its attached hydrogen atoms are written as a group. A subscript indicates the number of hydrogen atoms bonded to each carbon atom.

\[
\text{H} \\
| \text{H}--\text{C}-- \quad = \quad \text{CH}_3-- \\
| \text{H} \\
\text{Expanded} \quad \text{Condensed}
\]

When an organic molecule consists of a chain of three or more carbon atoms, the carbon atoms do not lie in a straight line. Rather, they are arranged in a zigzag pattern.

A simplified formula called the line-angle structural formula shows a zigzag line in which carbon atoms are represented as the ends of each line and as corners. For example, in the line-angle structural formula of pentane, each line in the zigzag drawing represents a single bond. The carbon atoms on the ends are bonded to three hydrogen atoms. However, the carbon atoms in the middle of the carbon chain are each bonded to two carbons and two hydrogen atoms as shown in Sample Problem 17.1.

SAMPLE PROBLEM 17.1 Drawing Expanded, Condensed, and Line-Angle Structural Formulas for an Alkane

Draw the expanded, condensed, and line-angle structural formulas for pentane.

TRY IT FIRST

SOLUTION

ANALYZE THE PROBLEM

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>pentane</td>
<td>structural formulas</td>
<td>carbon chain, H atoms, zigzag line</td>
</tr>
</tbody>
</table>

STEP 1 Draw the carbon chain. A molecule of pentane has five carbon atoms in a continuous chain.

C — C — C — C — C
17.1 Alkanes

**Guide to Drawing Structural Formulas for Alkanes**

**STEP 1** Draw the carbon chain.

**STEP 2** Draw the expanded structural formula by adding the hydrogen atoms using single bonds to each of the carbon atoms.

**STEP 3** Draw the condensed structural formula by combining the H atoms with each C atom.

**STEP 4** Draw the line-angle structural formula as a zigzag line in which the ends and corners represent C atoms.

---

**STEP 2** Draw the expanded structural formula by adding the hydrogen atoms using single bonds to each of the carbon atoms.

```
H H H H H
H—C—C—C—C—C—H
H H H H H
```

**STEP 3** Draw the condensed structural formula by combining the H atoms with each C atom.

```
H H H H H
H—C—C—C—C—C—H
CH₃—CH₂—CH₂—CH₂—CH₃
```

**STEP 4** Draw the line-angle structural formula as a zigzag line in which the ends and corners represent C atoms.

```
CH₃—CH₂—CH₂—CH₂—CH₃
```

---

**STUDY CHECK 17.1**

Draw the condensed structural formula and write the name for the following line-angle structural formula:

```
```

**ANSWER**

CH₃—CH₂—CH₂—CH₂—CH₂—CH₃  heptane

---

**Structural Isomers**

When an alkane has four or more carbon atoms, the atoms can be arranged so that a side group called a *branch* or *substituent* is attached to a carbon chain. For example, **FIGURE 17.2** shows two different ball-and-stick models for two compounds that have the molecular formula C₄H₁₀. One model is shown as a chain of four carbon atoms. In the other model, a carbon atom is attached as a branch or substituent to a carbon in a chain of three atoms. An alkane with at least one branch is called a *branched alkane*. When the two compounds have the same molecular formula but different arrangements of atoms, they are *structural isomers*.

**FIGURE 17.2** The structural isomers of C₄H₁₀ have the same number and type of atoms but are bonded in a different order.

What makes these molecules structural isomers?
In another example, we can draw the condensed and line-angle structural formulas for three different structural isomers with the molecular formula C₅H₁₂ as follows:

<table>
<thead>
<tr>
<th>Structural Isomers of C₅H₁₂</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Condensed</strong></td>
</tr>
<tr>
<td>CH₃—CH₂—CH₂—CH₂—CH₃</td>
</tr>
<tr>
<td>CH₃—CH—CH₂—CH₃</td>
</tr>
<tr>
<td>CH₃—C—CH₃</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 17.2 Structural Isomers**

Identify each pair of formulas as structural isomers or the same molecule.

a. CH₃—CH₂ and CH₂—CH₂—CH₃

b. and

**TRY IT FIRST**

**SOLUTION**

a. When we add up the number of C atoms and H atoms, they give the same molecular formula C₄H₁₀. Both of the structures are continuous four-carbon chains even though one or more of the —CH₃ ends are drawn above or below the horizontal part of the chain. Thus, both condensed structural formulas represent the same molecule and are not structural isomers.

b. When we add up the number of C atoms and H atoms, they give the same molecular formula C₆H₁₄. The line-angle structural formula on the left has a five-carbon chain with a —CH₃ substituent on the second carbon of the chain. The line-angle structural formula on the right has a four-carbon chain with two —CH₃ substituents. Thus, there is a different order of bonding of atoms, which represents structural isomers.

**STUDY CHECK 17.2**

Why does the following formula represent a different structural isomer of the molecules in Sample Problem 17.2, part b?

**ANSWER**

This formula represents a different structural isomer of C₆H₁₄ because the —CH₃ substituent is on the third carbon of the chain.

**Substituents in Alkanes**

In the IUPAC names for alkanes, a carbon branch is named as an alkyl group, which is an alkane that is missing one hydrogen atom. The alkyl group is named by replacing the ane ending of the corresponding alkane name with yl. Alkyl groups cannot exist on their own: They must be attached to a carbon chain. When a halogen atom is attached to a carbon
17.1 Alkanes

In the IUPAC system of naming, a carbon chain is numbered to give the location of the substituents. Let’s take a look at how we use the IUPAC system to name the alkane shown in Sample Problem 17.3.

**TABLE 17.3** Formulas and Names of Some Common Substituents

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₃</td>
<td>methyl</td>
</tr>
<tr>
<td>CH₃—CH₂—</td>
<td>ethyl</td>
</tr>
<tr>
<td>CH₃—CH—CH₃</td>
<td>propyl</td>
</tr>
<tr>
<td>CH₃—CH—CH₂—CH₂—</td>
<td>isopropyl</td>
</tr>
<tr>
<td>CH₃—CH—CH₂—CH₂—</td>
<td>isobutyl</td>
</tr>
<tr>
<td>CH₃—CH—CH₂—CH₃</td>
<td>sec-butyl</td>
</tr>
<tr>
<td>CH₃—C—CH₃</td>
<td>tert-butyl</td>
</tr>
<tr>
<td>F—Cl—</td>
<td>fluoro—chloro</td>
</tr>
<tr>
<td>Br—I—</td>
<td>bromo—iodo</td>
</tr>
</tbody>
</table>

**Naming Alkanes with Substituents**

In the IUPAC system of naming, a carbon chain is numbered to give the location of the substituents. Let’s take a look at how we use the IUPAC system to name the alkane shown in Sample Problem 17.3.

**SAMPLE PROBLEM 17.3** Writing IUPAC Names for Alkanes with Substituents

Write the IUPAC name for the following alkane:

CH₃ Br
CH₃—CH—CH₂—C—CH₂—CH₃

**TRY IT FIRST**

**SOLUTION**

**ANALYZE THE PROBLEM**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>six-carbon chain, two methyl groups, one bromo group</td>
<td>IUPAC name</td>
<td>position of substituents on the carbon chain</td>
</tr>
</tbody>
</table>

**STEP 1** Write the alkane name for the longest chain of carbon atoms.

CH₃ Br
CH₃—CH—CH₂—C—CH₂—CH₃
hexane

**STEP 2** Number the carbon atoms from the end nearer a substituent.

CH₃ Br
CH₃—CH—CH₂—C—CH₂—CH₃
hexane

1 2 3 4 5 6

**Guide to Naming Alkanes with Substituents**

**STEP 1**
Write the alkane name for the longest chain of carbon atoms.

**STEP 2**
Number the carbon atoms from the end nearer a substituent.

**STEP 3**
Give the location and name for each substituent (alphabetical order) as a prefix to the name of the main chain.
STEP Give the location and name for each substituent (alphabetical order) as a prefix to the name of the main chain. The substituents are listed in alphabetical order (bromo first, then methyl). A hyphen is placed between the number and the substituent name. When there are two or more of the same substituent, a prefix (di, tri, tetra) is used and commas separate the numbers. However, prefixes are not used to determine the alphabetical order of the substituents.

![Chemical structure](image)

4-bromo-2,4-dimethylhexane

STUDY CHECK 17.3

Write the IUPAC name for the following compound:

![Structural formula](image)

ANSWER

4-ethylheptane

Drawing Structural Formulas for Alkanes

The IUPAC name gives all the information needed to draw the condensed structural formula for an alkane. Suppose you are asked to draw the condensed structural formula for 2,3-dimethylbutane. The alkane name gives the number of carbon atoms in the longest chain. The names in the beginning indicate the substituents and where they are attached. We can break down the name in the following way:

<table>
<thead>
<tr>
<th>2,3-Dimethylbutane</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Substituents on</td>
<td></td>
</tr>
<tr>
<td>carbons 2 and 3</td>
<td></td>
</tr>
<tr>
<td>two identical</td>
<td></td>
</tr>
<tr>
<td>groups</td>
<td></td>
</tr>
<tr>
<td>—CH₃ alkyl groups</td>
<td></td>
</tr>
<tr>
<td>four C atoms in the</td>
<td></td>
</tr>
<tr>
<td>main chain</td>
<td></td>
</tr>
<tr>
<td>single (C — C) bonds</td>
<td></td>
</tr>
</tbody>
</table>

SAMPLE PROBLEM 17.4 Drawing Condensed Structures from IUPAC Names

Draw the condensed and line-angle structural formulas for 2,3-dimethylbutane.

TRY IT FIRST

SOLUTION

ANALYZE THE PROBLEM

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>2,3-dimethylbutane</td>
<td>condensed and line-angle structural formulas</td>
<td>four-carbon chain, two methyl groups</td>
</tr>
</tbody>
</table>

STEP 1 Draw the main chain of carbon atoms. For butane, we draw a chain of four carbon atoms or a zigzag line.

C —C —C —C
Guide to Drawing Structural Formulas for Alkanes with Substituents

STEP 1 Draw the main chain of carbon atoms.

STEP 2 Number the chain and place the substituents on the carbons indicated by the numbers.

STEP 3 For the condensed structural formula, add the correct number of hydrogen atoms to give four bonds to each C atom.

STUDY CHECK 17.4
Draw the condensed and line-angle structural formulas for 2-bromo-3-ethyl-4-methylpentane.

**ANSWER**

Properties and Uses of Alkanes

The first four alkanes—methane, ethane, propane, and butane—are gases at room temperature and are widely used as fuels.

Alkanes having five to eight carbon atoms (pentane, hexane, heptane, and octane) are liquids at room temperature. They are highly volatile, which makes them useful in fuels such as gasoline.

Liquid alkanes with 9 to 17 carbon atoms have higher boiling points and are found in kerosene, diesel, and jet fuels. Motor oil is a mixture of high-molecular-weight liquid hydrocarbons and is used to lubricate the internal components of engines. Mineral oil is a mixture of liquid hydrocarbons and is used as a laxative and a lubricant. Alkanes with 18 or more carbon atoms are waxy solids at room temperature. Known as paraffins, they are used in waxy coatings added to fruits and vegetables to retain moisture, inhibit mold, and enhance appearance. Petrolatum jelly, or Vaseline, is a semisolid mixture of hydrocarbons with more than 25 carbon atoms used in ointments and cosmetics and as a lubricant.

Reactions of Alkanes: Combustion

The carbon–carbon single bonds in alkanes are difficult to break, which makes them the least reactive family of organic compounds. However, alkanes burn readily (combustion) in oxygen to produce carbon dioxide, water, and energy.

\[
\text{Alkane}(g) + \text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + \text{H}_2\text{O}(g) + \text{energy}
\]
Propane is the gas used in portable heaters and gas barbecues (see FIGURE 17.3). The equation for the combustion of propane ($C_3H_8$) is written:

$$C_3H_8(g) + 5O_2(g) \xrightarrow{\Delta} 3CO_2(g) + 4H_2O(g) + \text{energy}$$

**Solubility and Density**

Alkanes are nonpolar, which makes them insoluble in water. However, they are soluble in nonpolar solvents. Alkanes have densities from 0.62 g/mL to about 0.79 g/mL, which is less than the density of water (1.0 g/mL).

If there is an oil spill in the ocean, the alkanes in the oil, which do not mix with water, form a thin layer on the surface that spreads over a large area (see FIGURE 17.4). In April 2010, an explosion on an oil-drilling rig in the Gulf of Mexico caused the largest oil spill in U.S. history. At its maximum, an estimated 10 million liters of oil was leaked every day. If the crude oil reaches land, there can be considerable damage to beaches, shellfish, birds, and wildlife habitats. When animals such as birds are covered with oil, they must be cleaned quickly because ingestion of the hydrocarbons when they try to clean themselves is fatal.

**FIGURE 17.3** The propane fuel in the tank undergoes combustion, which provides energy.

**FIGURE 17.4** In oil spills, large quantities of oil spread out to form a thin layer on top of the ocean surface.

**QUESTIONS AND PROBLEMS**

### 17.1 Alkanes

**LEARNING GOAL** Write the IUPAC names and draw the condensed or line-angle structural formulas for alkanes.

17.1 Write the IUPAC name for each of the following alkanes:

- a. $\text{CH}_3$ $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_3$
- b. $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_3$
- c. $\text{CH}_3$ $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_3$

17.2 Write the IUPAC name for each of the following alkanes:

- a. $\text{CH}_3$ $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_3$
- b. $\text{CH}_3$ $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_3$
- c. $\text{CH}_3$ $\text{CH}_2$ $\text{CH}_2$ $\text{CH}_3$

17.3 Indicate whether each of the following pairs of formulas represents structural isomers or the same molecule:

- a. $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$ and $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$
- b. $\text{CH}_2$ $\text{CH}$ $\text{CH}_2$ $\text{CH}_3$ and $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$
- c. $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$ and $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$

17.4 Indicate whether each of the following pairs of formulas represents structural isomers or the same molecule:

- a. $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$ and $\text{CH}$ $\text{CH}_2$ $\text{CH}_3$
- b. $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$ and $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$
- c. $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$ and $\text{CH}_3$ $\text{CH}$ $\text{CH}_3$
17.2 Alkenes, Alkynes, and Polymers

**LEARNING GOAL** Write the IUPAC names and draw the condensed or line-angle structural formulas for alkenes and alkynes.

We organize organic compounds into *classes* or *families* by their *functional groups*, which are groups of specific atoms. Compounds that contain the same functional group have similar physical and chemical properties. Identifying functional groups allows us to classify organic compounds according to their structure, to name compounds within each family, predict their chemical reactions, and draw the structures for their products.

Alkenes and alkynes are classes of unsaturated hydrocarbons that contain *double* and *triple* bonds, respectively. The functional group of an *alkene* contains at least one double bond between carbons. The double bond forms when two adjacent carbon atoms share two pairs of valence electrons. The simplest alkene is ethene, C₂H₄, which is often called by its common name, ethylene. In ethene, each carbon atom is attached to two H atoms and the other carbon atom in the double bond. The resulting molecule has a flat geometry because the carbon and hydrogen atoms all lie in the same plane. The functional group of an *alkyne* contains a triple bond, which occurs when two carbon atoms share three pairs of valence electrons.
A list of the common functional groups in organic compounds is shown in **Table 17.4**.

**Table 17.4 Classes of Organic Compounds**

<table>
<thead>
<tr>
<th>Class</th>
<th>Functional Group</th>
<th>Example/Name</th>
<th>Occurrence</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alkene</td>
<td>( \overset{\bigwedge}{\overset{\wedge}{\bigwedge}} C \overset{\bigwedge}{\wedge} C )</td>
<td>( H_2C=CH_2 ) Ethene (ethylene)</td>
<td>Ripening of fruit, used to make polyethylene</td>
</tr>
<tr>
<td>Alkyne</td>
<td>( \overset{\wedge}{\bigwedge} C \overset{\wedge}{\overset{\bigwedge}{\wedge}} C \overset{\wedge}{\bigwedge} C )</td>
<td>( H-C=CH ) Ethyne (acetylene)</td>
<td>Welding fuel</td>
</tr>
<tr>
<td>Alcohol</td>
<td>( -OH )</td>
<td>( CH_3=CH_2-OH ) Ethanol (ethyl alcohol)</td>
<td>Solvent</td>
</tr>
<tr>
<td>Ether</td>
<td>( -O- )</td>
<td>( CH_3-C=O-CH_2=CH_3 ) Diethyl ether</td>
<td>Solvent</td>
</tr>
<tr>
<td>Aldehyde</td>
<td>( \overset{\wedge}{\bigwedge} O )</td>
<td>( CH_3-C=H ) Ethanal (acetaldehyde)</td>
<td>Preparation of acetic acid</td>
</tr>
<tr>
<td>Ketone</td>
<td>( \overset{\wedge}{\bigwedge} O )</td>
<td>( CH_3-C=CH_3 ) Propanone (acetone)</td>
<td>Solvent, paint and fingernail polish remover</td>
</tr>
<tr>
<td>Carboxylic acid</td>
<td>( \overset{\wedge}{\bigwedge} O )</td>
<td>( CH_3-C=OH ) Ethanoic acid (acetic acid)</td>
<td>Component of vinegar</td>
</tr>
<tr>
<td>Ester</td>
<td>( \overset{\wedge}{\bigwedge} O )</td>
<td>( CH_3-C=O-CH_3 ) Methyl ethanoate (methyl acetate)</td>
<td>Rum flavoring</td>
</tr>
<tr>
<td>Amine</td>
<td>( \overset{\wedge}{\bigwedge} N )</td>
<td>( CH_3-NH_2 ) Methylamine</td>
<td>Odor of fish</td>
</tr>
<tr>
<td>Amide</td>
<td>( \overset{\wedge}{\bigwedge} O )</td>
<td>( CH_3-C-NH_2 ) Ethanamide (acetamide)</td>
<td>Odor of mice</td>
</tr>
</tbody>
</table>

**Naming Alkenes and Alkynes**

The IUPAC names for alkenes and alkynes are similar to those of alkanes. The IUPAC name of the simplest alkyne is ethyne, although it is often called by the common name, acetylene (see **Table 17.5**). When naming alkenes and alkynes, the longest carbon chain must contain the double or triple bond.

**Table 17.5 Comparison of Alkanes, Alkenes, and Alkynes**

<table>
<thead>
<tr>
<th>Alkane</th>
<th>Alkene</th>
<th>Alkyne</th>
</tr>
</thead>
<tbody>
<tr>
<td>( CH_3=CH_3 )</td>
<td>( H_2C=CH_2 )</td>
<td>( H-C=CH ) Ethyne (acetylene)</td>
</tr>
<tr>
<td>Ethane</td>
<td></td>
<td></td>
</tr>
<tr>
<td>( CH_3=CH_2=CH_3 )</td>
<td>( CH_3=CH=CH_2 )</td>
<td>( CH_3=C=CH )丙yne (propyne)</td>
</tr>
<tr>
<td>Propane</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Fruit is ripened with ethene, a plant hormone.

A mixture of acetylene and oxygen undergoes combustion during the welding of metals.
An example of naming an alkene is shown in Sample Problem 17.5.

**SAMPLE PROBLEM 17.5 Naming Alkenes and Alkynes**
Write the IUPAC name for the following:
\[
\text{CH}_3 \\
\text{CH}_3 - \text{CH} - \text{CH} = \text{CH} - \text{CH}_3
\]

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>five-carbon chain,</td>
<td>IUPAC name</td>
<td>replace the ane in the</td>
</tr>
<tr>
<td></td>
<td>double bond, methyl</td>
<td></td>
<td>alkane name with ene</td>
</tr>
<tr>
<td></td>
<td>group</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STEP 1** Name the longest carbon chain that contains the double bond. There are five carbon atoms in the longest carbon chain containing the double bond. Replacing the corresponding alkane ending with *ene* gives pentene.

\[
\text{CH}_3 \\
\text{CH}_3 - \text{CH} - \text{CH} = \text{CH} - \text{CH}_3 \\
\text{pentene}
\]

**STEP 2** Number the carbon chain from the end nearer the double bond. The number of the first carbon in the double bond is used to give the location of the double bond. Alkenes or alkynes with two or three carbons do not need numbers.

\[
\text{CH}_3 \\
\text{CH}_3 - \text{CH} - \text{CH} = \text{CH} - \text{CH}_3 \\
\text{2-pentene}
\]

**STEP 3** Give the location and name for each substituent (alphabetical order) as a prefix to the alkene name. The methyl group is located on carbon 4.

\[
\text{CH}_3 \\
\text{CH}_3 - \text{CH} - \text{CH} = \text{CH} - \text{CH}_3 \\
\text{4-methyl-2-pentene}
\]

**STUDY CHECK 17.5**
Write the IUPAC name for the following:
\[
\text{CH}_3 - \text{CH}_2 - \text{C} = \text{C} - \text{CH}_2 - \text{CH}_3
\]

**ANSWER**
3-hexyne

**Hydrogenation: Addition Reaction**
In a reaction called *hydrogenation*, H atoms add to each of the carbon atoms in a double bond of an alkene. During hydrogenation, the double bonds are converted to single bonds in alkanes. A metal catalyst such as finely divided platinum (Pt), nickel (Ni), or palladium (Pd) is used to speed up the reaction.

\[
\text{CH}_3 - \text{CH} = \text{CH} - \text{CH}_3 + \text{H}_2 \xrightarrow{\text{Pt}} \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_3
\]

**Guide to Naming Alkenes and Alkynes**

**STEP 1** Name the longest carbon chain that contains the double or triple bond.

**STEP 2** Number the carbon chain from the end nearer the double or triple bond.

**STEP 3** Give the location and name for each substituent (alphabetical order) as a prefix to the alkene or alkyne name.
SAMPLE PROBLEM 17.6 Writing an Equation for Hydrogenation

Draw the condensed structural formula for the product of the following hydrogenation reaction:

\[
\text{CH}_3\text{CH}=\text{CH}_2 + \text{H}_2 \xrightarrow{\text{Pt}} \text{CH}_3\text{CH}_2\text{CH}_3
\]

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>alkene + H₂</td>
<td>structural formula of product</td>
<td>add 2H to double bond</td>
<td></td>
</tr>
</tbody>
</table>

\[
\text{CH}_3\text{CH} = \text{CH}_2 \xrightarrow{\text{Pt}} \text{CH}_3\text{CH}_2\text{CH}_3
\]

STUDY CHECK 17.6

Draw the condensed structural formula for the product of the hydrogenation of 2-methyl-1-butene, using a platinum catalyst.

ANSWER

\[
\begin{align*}
\text{CH}_3 \\
\text{CH}_3\text{CH} \xrightarrow{\text{Pt}} \text{CH}_2\text{CH}_3
\end{align*}
\]

CHEMISTRY LINK TO HEALTH

Hydrogenation of Unsaturated Fats

Vegetable oils such as corn oil or safflower oil are unsaturated fats composed of fatty acids that contain double bonds. The process of hydrogenation is used commercially to convert the double bonds in the unsaturated fats in vegetable oils to saturated fats such as margarine, which are more solid. Adjusting the amount of added hydrogen produces partially hydrogenated fats such as soft margarine, solid margarine in sticks, and shortenings, which are used in cooking. For example, oleic acid is a typical unsaturated fatty acid in olive oil and has a double bond at carbon 9. When oleic acid is hydrogenated, it is converted to stearic acid, a saturated fatty acid.

Polymers

A polymer is a large molecule that consists of small repeating units called monomers. In the past hundred years, the plastics industry has made synthetic polymers that are in many of the materials we use every day, such as carpeting, plastic wrap, nonstick pans, plastic cups, and rain gear (see FIGURE 17.5). In medicine, synthetic polymers are used to replace...
diseased or damaged body parts such as hip joints, teeth, heart valves, and blood vessels. There are about 300 billion kg of plastics produced every year, which is about 40 kg for every person on Earth.

Many of the synthetic polymers are made by reactions of small alkene monomers. Polymerization reactions require high temperature, a catalyst, and high pressure (over 1000 atm). A polymer grows longer as each monomer is added at the end of the chain and contains as many as 1000 monomers. Polyethylene, a polymer made from ethylene monomers, is used in plastic bottles, film, and plastic dinnerware. More polyethylene is produced worldwide than any other polymer. Low-density polyethylene (LDPE) is flexible, breakable, less dense, and more branched than high-density polyethylene (HDPE). High-density polyethylene is stronger, more dense, and melts at a higher temperature than LDPE.

**TABLE 17.6** lists several alkene monomers that are used to produce common synthetic polymers. The alkane-like nature of these plastic synthetic polymers makes them unreactive. Thus, they do not decompose easily (they are not biodegradable). As a result, they have become significant contributors to pollution, on land and in the oceans. Efforts are being made to make them more degradable.

You can identify the type of polymer used to manufacture a plastic item by looking for the recycling symbol (arrows in a triangle) found on the label or on the bottom of the plastic container. For example, the number 5 or the letters PP inside the triangle is the code for a polypropylene plastic. There are now many cities that maintain recycling programs that reduce the amount of plastic materials that are transported to landfills.

Today, products such as lumber, tables and benches, trash receptacles, and pipes used for irrigation systems are made from recycled plastics.
Chapter 17
Organic Chemistry

Sample Problem 17.7 Polymers

A firefighter/EMT arrives at a home where a premature baby has been delivered. To prevent hypothermia during transport to the neonatal facility, she wraps the baby in cling wrap, which is polydichloroethylene. Draw the monomer unit, which is 1,1-dichloroethene, and draw a portion of the polymer formed from three monomer units.

Try It First

Solution

\[
\begin{align*}
\text{Cl} & \quad \text{Cl} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\end{align*}
\]

1,1-Dichloroethene

\[
\begin{align*}
\text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} \\
\text{H} & \quad \text{H} & \quad \text{H} & \quad \text{H} & \quad \text{H} & \quad \text{H} \\
\text{Cl} & \quad \text{Cl} & \quad \text{Cl} & \quad \text{Cl} & \quad \text{Cl} & \quad \text{Cl} \\
\text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} \\
\text{Cl} & \quad \text{Cl} & \quad \text{Cl} & \quad \text{Cl} & \quad \text{Cl} & \quad \text{Cl} \\
\end{align*}
\]

Polydichloroethylene

---

Table 17.6 Some Alkenes and Their Polymers

<table>
<thead>
<tr>
<th>Recycling Code</th>
<th>Monomer</th>
<th>Polymer Section</th>
<th>Common Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Ethene (ethylene)</td>
<td></td>
<td>Plastic bottles, film, insulation materials, shopping bags</td>
</tr>
<tr>
<td>4</td>
<td>HDPE</td>
<td></td>
<td>Plastic pipes and tubing, garden hoses, garbage bags, shower curtains, credit cards, medical containers</td>
</tr>
<tr>
<td>3</td>
<td>Chloroethene</td>
<td></td>
<td>Polyvinyl chloride (PVC)</td>
</tr>
<tr>
<td>5</td>
<td>Propene</td>
<td></td>
<td>Polypropylene (PP)</td>
</tr>
<tr>
<td>6</td>
<td>Polystyrene</td>
<td></td>
<td>Plastic coffee cups and cartons, insulation</td>
</tr>
</tbody>
</table>

---

F"
STUDY CHECK 17.7
Draw the condensed structural formula for the monomer used in the manufacturing of PVC.

ANSWER
\[
\text{Cl} \quad \text{H}_2\text{C} \equiv \text{CH}
\]

QUESTIONS AND PROBLEMS

17.2 Alkenes, Alkynes, and Polymers

LEARNING GOAL Write the IUPAC names and draw the condensed or line-angle structural formulas for alkenes and alkynes.

17.11 Identify each of the following as an alkane, alkene, or alkyne:
   a. \(\text{CH}_3 \text{CH} \equiv \text{CH}_2\)
   b. \(\text{CH}_3 \text{CH}_2 \text{C} \equiv \text{CH}\)
   c. 

17.12 Identify each of the following as an alkane, alkene, or alkyne:
   a. 
   b. \(\text{CH}_3 \text{C} \equiv \text{C} \text{CH}_3\)
   b. \(\text{CH}_3 \text{CHCH}_2 \text{CH}_3\)
   c. \(\text{CH}_3 \text{CH}_2 \text{CH} \equiv \text{CH}\)

17.13 Write the IUPAC name for each of the following:
   a. \(\text{H}_2\text{C} \equiv \text{CH}_2\)
   b. \(\text{CH}_3 \text{C} \equiv \text{C} \text{CH}_3\)
   c. \(\text{CH}_3 \text{CH}_2 \text{C} \equiv \text{C} \text{CH}_3\)

17.14 Write the IUPAC name for each of the following:
   a. \(\text{H}_2\text{C} \equiv \text{CH} \equiv \text{CH}_2 \text{CH}_3\)
   b. \(\text{CH}_3 \equiv \text{C} \equiv \text{C} \equiv \text{C} \equiv \text{C} \text{CH}_3\)
   c. \(\text{CH}_3 \text{CH}_2 \text{C} \equiv \text{CH} \equiv \text{CH}_3\)

17.15 Draw the condensed structural formula for each of the following compounds:
   a. propene
   b. 1-hexyne
   c. 2-methyl-1-butene

17.16 Draw the condensed structural formula for each of the following compounds:
   a. \(\text{CH}_3 \equiv \text{CHCH}_2 \text{CH}_2 \text{CH}_2 \text{H}_2\text{CH}_3\)
   b. \(\text{CH}_3 \equiv \text{CHCH}_2 \text{CH}_3\)
   c. \(\text{CH}_3 \equiv \text{CHCH}_2 \text{CH}_3\)

17.17 Draw the condensed structural formula for the product in each of the following:
   a. \(\text{CH}_3 \equiv \text{CHCH}_2 \text{CH}_2 \text{CH}_2 \equiv \text{CH}_2 + \text{H}_2 \text{Pt}\rightarrow\)
   b. \(\text{CH}_3 \equiv \text{CHCH}_2 \equiv \text{CH} + \text{H}_2 \text{Ni}\rightarrow\)
   c. 

17.18 Draw the condensed structural formula for the product in each of the following:
   a. \(\text{CH}_3 \equiv \text{CHCH}_2 \equiv \text{CH} + \text{H}_2 \text{Pt}\rightarrow\)
   b. 
   c. 

17.19 What is a polymer?

17.20 What is a monomer?

17.21 Write an equation that represents the formation of a portion of polypropylene from three of its monomers.

17.22 Write an equation that represents the formation of a portion of polystyrene from three of its monomers.

Applications

17.23 The plastic polyvinylidene difluoride, PVDF, is made from monomers of 1,1-difluoroethene. Draw the expanded structural formula for a portion of the polymer formed from three monomers of 1,1-difluoroethene.

17.24 The polymer polyacrylonitrile, PAN, used in the fabric material Orlon is made from monomers of acrylonitrile. Draw the expanded structural formula for a portion of the polymer formed from three monomers of acrylonitrile.
17.3 Aromatic Compounds

LEARNING GOAL Describe the bonding in benzene; name aromatic compounds, and draw their line-angle structural formulas.

In 1825, Michael Faraday isolated a hydrocarbon called benzene, which had the molecular formula $C_6H_6$. A molecule of benzene consists of a ring of six carbon atoms with one hydrogen atom attached to each carbon. Because many compounds containing benzene had fragrant odors, the family of benzene compounds became known as aromatic compounds.

In benzene, each carbon atom uses three valence electrons to bond to the hydrogen atom and two adjacent carbons. That leaves one valence electron, which scientists first thought was shared in a double bond with an adjacent carbon. In 1865, August Kekulé proposed that the carbon atoms in benzene were arranged in a flat ring with alternating single and double bonds between the carbon atoms. There are two possible structural representations of benzene due to resonance in which the double bonds can form between two different carbon atoms.

Today, we know that the six electrons are shared equally among the six carbon atoms and that all the carbon–carbon bonds in benzene are identical due to resonance. This unique feature of benzene makes it especially stable. Benzene is most often represented as a line-angle structural formula, which shows a hexagon with a circle in the center. Some of the ways to represent benzene are shown as follows:

![Structural representations for benzene](image)

Some common examples of aromatic compounds that we use for flavoring are anisole from anise, estragole from tarragon, and thymol from thyme.

![Anise (anisole), Tarragon (estragole), Thyme (thymol)](image)

The aroma and flavor of the herbs anise, tarragon, and thyme are because of aromatic compounds.

Naming Aromatic Compounds

Many compounds containing benzene have been important in chemistry for many years and still use their common names. Names such as toluene, aniline, and phenol are allowed by IUPAC rules. When benzene has only one substituent, the ring is not numbered.
When there are two substituents, the benzene ring is numbered to give the lowest numbers to the substituents. When a common name can be used such as toluene, aniline, or phenol, the carbon atom attached to the methyl, amine, or hydroxyl group, is numbered as carbon 1.

How is the structure of toluene similar to and different from that of phenol?

Sample Problem 17.8 Naming Aromatic Compounds

Write the IUPAC name for the following compound:

\[
\text{Br} \quad \text{NH}_2
\]

Try It First

Solution

Step 1 Write the name for the aromatic compound. This aromatic compound contains a benzene ring with an amine group, which is aniline.
**Guide to Naming Aromatic Compounds**

**STEP 1**
Write the name for the aromatic compound.

**STEP 2**
If there are two substituents, number the aromatic ring.

**STEP 3**
Name a substituent as a prefix.

---

**STEP 2** If there are two substituents, number the aromatic ring. The benzene ring is numbered to give the lower number when counting the carbon with the amine group as 1. The bromine atom is on carbon 2.

**STEP 3** Name a substituent as a prefix.

\[
\begin{align*}
\text{NH}_2 & \quad \text{Br} \\
& \quad 2\text{-bromoaniline}
\end{align*}
\]

**STUDY CHECK 17.8**
Write the IUPAC name for the following compound:

\[
\begin{align*}
\text{CH}_3 \\
& \quad \text{Cl}
\end{align*}
\]

**ANSWER**
3-chlorotoluene

---

**QUESTIONS AND PROBLEMS**

**17.3 Aromatic Compounds**

**LEARNING GOAL** Describe the bonding in benzene; name aromatic compounds, and draw their line-angle structural formulas.

17.25 Write the IUPAC name for each of the following:

a. \[
\begin{align*}
\text{CH}_3 \\
& \quad \text{Cl} \\
& \quad \text{CH}_2 \quad \text{CH}_3
\end{align*}
\]

b. \[
\begin{align*}
\text{OH}
\end{align*}
\]

c. \[
\begin{align*}
\text{Cl}
\end{align*}
\]

**17.26** Write the IUPAC name for each of the following:

a. \[
\begin{align*}
\quad \text{Br}
\end{align*}
\]

b. \[
\begin{align*}
\quad \text{CH}_3
\end{align*}
\]

c. \[
\begin{align*}
\quad \text{Cl}
\end{align*}
\]

**17.27** Draw the line-angle structural formula for each of the following compounds:

a. aniline
b. 1-chloro-4-fluorobenzene
c. 4-ethyltoluene

**17.28** Draw the line-angle structural formula for each of the following compounds:

a. 1,3-dibromo-5-chlorobenzene
b. 2,4-dichlorotoluene
c. propylbenzene

---

**17.4 Alcohols and Ethers**

**LEARNING GOAL** Write the IUPAC and common names for alcohols and ethers; draw the condensed or line-angle structural formulas when given their names.

In the functional group of an **alcohol**, a hydroxyl group (—OH) replaces a hydrogen atom in a hydrocarbon. In a phenol, the hydroxyl group replaces a hydrogen atom attached to a benzene ring. Molecules of alcohols and phenols have a bent shape around the oxygen atom. In the functional group of an **ether**, an oxygen atom is attached by single bonds to two carbon atoms (see **FIGURE 17.6**).
**Naming Alcohols**

In the IUPAC system, an alcohol is named by replacing the final *e* in the corresponding alkane name with *ol*. The common name of a simple alcohol uses the name of the alkyl group followed by *alcohol*.

![Structural formulas of alcohols](image)

Alcohols with one or two carbon atoms do not require a number for the hydroxyl group. When an alcohol consists of a chain with three or more carbon atoms, the chain is numbered to give the position of the —OH group and any substituents on the chain.

![Alcohols with numbered carbon chains](image)

We can also draw the line-angle structural formulas for alcohols as shown for 2-propanol and 2-butanol.

**SAMPLE PROBLEM 17.9 Naming Alcohols**

Write the IUPAC name for the following:

![Alcohol structure](image)

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>alcohol, five-carbon chain, methyl substituent</td>
<td>IUPAC name</td>
<td>position of —CH₃ and —OH groups, replace <em>e</em> in alkane name with <em>ol</em></td>
</tr>
</tbody>
</table>
**Guide to Naming Alcohols**

**STEP 1** Name the longest carbon chain attached to the \( -\text{OH} \) group by replacing the \( e \) in the corresponding alkane name with \( ol \). Name an aromatic alcohol as a **phenol**.

**STEP 2** Number the chain from the end nearer the \( -\text{OH} \) group.

**STEP 3** Give the location and name for each substituent relative to the \( -\text{OH} \) group.

**STEP** Name the longest carbon chain attached to the \( -\text{OH} \) group by replacing the \( e \) in the corresponding alkane name with \( ol \). To name the alcohol, the \( e \) in alkane name pentane is replaced by \( ol \).

\[
\begin{align*}
\text{CH}_3 & \quad \text{OH} \\
\text{CH}_3 \quad \text{CH} & \quad \text{CH}_2 \quad \text{CH} \quad \text{CH} \\
\text{pentanol}
\end{align*}
\]

**STEP** Number the chain from the end nearer the \( -\text{OH} \) group. This carbon chain is numbered from right to left to give the position of the \( -\text{OH} \) group as carbon 2, which is shown as a prefix in the name 2-pentanol.

\[
\begin{align*}
5 \quad 4 \quad 3 \quad 2 \quad 1 \\
\text{CH}_3 \quad \text{CH} & \quad \text{CH}_2 \quad \text{CH} \quad \text{CH} \\
\text{2-pentanol}
\end{align*}
\]

**STEP** Give the location and name of each substituent relative to the \( -\text{OH} \) group.

\[
\begin{align*}
5 \quad 4 \quad 3 \quad 2 \quad 1 \\
\text{CH}_3 \quad \text{CH} & \quad \text{CH}_2 \quad \text{CH} \quad \text{CH} \\
\text{4-methyl-2-pentanol}
\end{align*}
\]

**STUDY CHECK 17.9**

Write the IUPAC name for the following:

\[
\begin{align*}
\text{Cl} & \quad \text{OH} \\
\text{3-chloro-1-butanol}
\end{align*}
\]

**ANSWER**

3-chloro-1-butanol

**Naming Ethers**

An **ether** consists of an oxygen atom that is attached by single bonds to two carbon groups that are alkyl or aromatic groups. In the common name of an ether, the names of the alkyl or aromatic groups attached to the oxygen atom are written in alphabetical order, followed by the word **ether**.

\[
\begin{align*}
\text{Methyl group} & \quad \text{Propyl group} \\
\text{CH}_3 & \quad \text{O} \quad \text{CH}_3 \quad \text{CH}_2 \quad \text{CH}_3 \\
\text{Common name: methyl propyl ether}
\end{align*}
\]

**CHEMISTRY LINK TO HEALTH**

Some Important Alcohols, Phenols, and Ethers

**Methanol** (methyl alcohol), the simplest alcohol, is found in solvents and paint removers. If ingested, methanol is oxidized to formaldehyde, which can cause headaches, blindness, and death. Methanol is used to make plastics, medicines, and fuels. In car racing, it is used as a fuel because it is less flammable and has a higher octane rating than does gasoline.

**Ethanol** (ethyl alcohol) is used in hand sanitizers. In an alcohol-containing sanitizer, the amount of ethanol is typically 60% (v/v) but can be as high as 85% (v/v). This amount of ethanol can make hand sanitizers a fire hazard in the home because ethanol is highly flammable. When using an ethanol-containing sanitizer, it is important to rub hands until they are completely dry. It is also recommended that sanitizers containing ethanol be placed in storage areas that are away from heat sources in the home.

**1,2,3-Propanetriol** (glycerol or glycerin), a trihydroxy alcohol, is a viscous liquid obtained from oils and fats during the production of soaps. The presence of several polar \( -\text{OH} \) groups makes it strongly attracted to water, a feature that makes glycerin useful as a skin softener in products such as skin lotions, cosmetics, shaving creams, and liquid soaps.
1,2-Ethanediol (ethylene glycol) is used as an antifreeze in heating and cooling systems. It is also a solvent for paints, inks, and plastics, and it is used in the production of synthetic fibers such as Dacron. If ingested, it is extremely toxic. In the body, it is oxidized to oxalic acid, which forms insoluble compounds in the kidneys that cause renal damage, convulsions, and death. Because its sweet taste is attractive to pets and children, ethylene glycol solutions must be carefully stored.

Phenols are found in several of the essential oils of plants, which produce the odor or flavor of the plant. Eugenol is found in cloves, vanillin in vanilla bean, isoeugenol in nutmeg, and thymol in thyme and mint. Thymol has a pleasant, minty taste and is used in mouthwashes and by dentists to disinfect a cavity before adding a filling compound.

Anesthesia is the loss of sensation and consciousness. The term ether has been associated with anesthesia because diethyl ether was the most widely used anesthetic for more than a hundred years. Although it is easy to administer, ether is very volatile and highly flammable. A small spark in the operating room could cause an explosion. Since the 1950s, anesthetics such as Forane (isoflurane), Ethrane (enflurane), Suprane (desflurane), and Sevoflurane have been developed that are not as flammable. These anesthetics retain the ether group, but the addition of halogen atoms reduces the volatility and flammability of the ethers.
17.5 Aldehydes and Ketones

**LEARNING GOAL** Write the IUPAC and common names for aldehydes and ketones; draw the condensed or line-angle structural formulas when given their names.

The carbonyl group (C=O) has a carbon–oxygen double bond with two groups of atoms attached to the carbon atom at angles of 120°. Because the oxygen atom in the carbonyl group is much more electronegative than the carbon atom, the carbonyl group has a dipole with a partial negative charge (δ⁻) on the oxygen and a partial positive charge (δ⁺) on the carbon.

In the functional group of an aldehyde, the carbonyl group is bonded to at least one hydrogen atom. That carbon may also be bonded to another hydrogen, a carbon atom in an alkyl group, or an aromatic ring (see **FIGURE 17.7**). In the functional group of a ketone, the carbonyl group is bonded to two carbon atoms.

**FIGURE 17.7** The carbonyl group is found in aldehydes and ketones.

Q If aldehydes and ketones both contain a carbonyl group, how can you differentiate between compounds from each family?
Naming Aldehydes

In the IUPAC system, an aldehyde is named by replacing the e in the corresponding alkane name with al. No number is needed for the aldehyde group because it always appears at the end of the chain. The aldehydes with carbon chains of one to four carbon atoms are often referred to by their common names, which end in aldehyde. The roots (form, acet, propion, and butyr) of these common names are derived from Latin or Greek words (see FIGURE 17.8).

The carbonyl carbon is at the end of the chain

FIGURE 17.8 In the structures of aldehydes, the carbonyl group is always the end carbon.

Why is the carbon in the carbonyl group in aldehydes always at the end of the chain?

The aldehyde of benzene is named benzaldehyde.

The carbonyl carbon is at the end of the chain

SAMPLE PROBLEM 17.10 Naming Aldehydes

Write the IUPAC name for the following:

CH₃
CH₂CH₂CHCH₂C

TRY IT FIRST

SOLUTION

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>five-carbon chain,</td>
<td></td>
<td>IUPAC name</td>
<td>position of —CH₃, methyl substituent replace e in alkane name with al</td>
</tr>
</tbody>
</table>
**Guide to Naming Aldehydes**

**STEP 1** Name the longest carbon chain by replacing the *e* in the alkane name with *al*. The longest carbon chain containing the carbonyl group has five carbon atoms.

\[
\begin{align*}
\text{CH}_3 & \quad \text{O} \\
\text{CH}_3 & \quad \text{CH}_2 & \quad \text{CH} & \quad \text{CH}_2 & \quad \text{C} & \quad \text{H} \\
\text{pentanal}
\end{align*}
\]

**STEP 2** Name and number any substituents by counting the carbonyl group as carbon 1. Counting from the right, the methyl substituent is on carbon 3.

\[
\begin{align*}
\text{CH}_3 & \quad \text{O} \\
\text{CH}_3 & \quad \text{CH}_2 & \quad \text{CH} & \quad \text{CH}_2 & \quad \text{C} & \quad \text{H} \\
\text{3-methylpentanal}
\end{align*}
\]

**STUDY CHECK 17.10**

What are the IUPAC and common names of the aldehyde with three carbon atoms?

**ANSWER**

propanal (propionaldehyde)

---

**CHEMISTRY LINK TO THE ENVIRONMENT**

**Vanilla**

Vanilla has been used as a flavoring for over a thousand years. After drinking a beverage made from powdered vanilla and cocoa beans with Emperor Montezuma in Mexico, the Spanish conquistador Hernán Cortés took vanilla back to Europe where it became popular for flavoring and for scenting perfumes and tobacco. Thomas Jefferson introduced vanilla to the United States during the late 1700s. Today much of the vanilla we use in the world is grown in Indonesia, Madagascar, Mexico, Papua New Guinea, and China.

The vanilla plant is a member of the orchid family and thrives under tropical conditions. There are many species of Vanilla, but **Vanilla planifolia** (or **V. fragrans**) is considered to produce the best flavor. The vanilla plant grows like a vine, which can attain a length of 100 feet. Its flowers are hand-pollinated to produce a green fruit that is picked in eight or nine months. The fruit is sun-dried so that it becomes a long, dark brown pod, which is called a “vanilla bean” because it looks like a string bean. The flavor and fragrance of the vanilla bean comes from the tiny black seeds found inside the dried bean.

The seeds and pod are used to flavor desserts such as custards and ice cream. The extract of vanilla is made by chopping up vanilla beans and mixing them with a 35% ethanol–water mixture. The liquid, which contains the aldehyde **vanillin**, is drained from the bean residue and used for flavoring.

The vanilla bean is the dried fruit of the vanilla plant.

Vanilla flavoring is prepared by soaking vanilla beans in ethanol and water.
Naming Ketones

In the IUPAC system, the name of a ketone is obtained by replacing the \( e \) in the corresponding alkane name with \( one \). However, the common names for unbranched ketones are still in use. In the common names, the alkyl groups bonded on either side of the carbonyl group are listed alphabetically, followed by \( ketone \). Acetone, which is another name for propanone, has been retained by the IUPAC system. The naming of a ketone is shown in Sample Problem 17.11.

**SAMPLE PROBLEM 17.11 Naming Ketones**

Write the IUPAC name for the following ketone:

\[
\begin{array}{c}
\text{O} \\
\text{\_\_\_\_\_\_\_} \\
\text{\_\_\_\_\_\_\_} \\
\end{array}
\]

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>Name the longest carbon chain by replacing the ( e ) in the alkane name with ( one ).</td>
<td>five-carbon chain, methyl substituent</td>
<td>IUPAC name</td>
<td>position of (-\text{CH}_2) and carbonyl, replace ( e ) in alkane name with ( one )</td>
</tr>
</tbody>
</table>

**STEP 1** Name the longest carbon chain by replacing the \( e \) in the alkane name with \( one \).

\[
\begin{array}{c}
\text{O} \\
\text{\_\_\_\_\_\_\_} \\
\text{\_\_\_\_\_\_\_} \\
\end{array}
\]

pentanone

**STEP 2** Number the carbon chain from the end nearer the carbonyl group and indicate its location.

\[
\begin{array}{c}
\text{O} \\
\text{\_\_\_\_\_\_\_} \\
\text{\_\_\_\_\_\_\_} \\
\end{array}
\]

2-pentanone

**STEP 3** Name and number any substituents on the carbon chain. Counting from the right, the methyl substituent is on carbon 4.

\[
\begin{array}{c}
\text{O} \\
\text{\_\_\_\_\_\_\_} \\
\text{\_\_\_\_\_\_\_} \\
\end{array}
\]

4-methyl-2-pentanone

**STUDY CHECK 17.11**

What is the common name of 3-hexanone?

**Answer**

ethyl propyl ketone
## 17.5 Aldehydes and Ketones

**LEARNING GOAL** Write the IUPAC and common names for aldehydes and ketones; draw the condensed or line-angle structural formulas when given their names.

**17.35** Write the common name for each of the following:

- \( CH_3 - C \rightarrow H \)
- \( CH_3 - CH_2 - C \rightarrow CH_3 \)
- \( H \rightarrow C \rightarrow H \)

**17.36** Write the common name for each of the following:

- \( CH_3 - CH_2 - C \rightarrow CH_3 \)
- \( CH_3 \)
- \( CH_3 - CH_2 - C \rightarrow H \)

**17.37** Write the IUPAC name for each of the following:

- \( CH_3 - CH_2 - C \rightarrow H \)

**17.38** Write the IUPAC name for each of the following:

- \( CH_3 - CH_2 - CH_2 - C \rightarrow H \)
- \( CH_3 \)
- \( C \rightarrow H \)

**17.39** Draw the condensed structural formula for each of the following:

- acetaldehyde
- 2-pentanone
- butyl methyl ketone

**17.40** Draw the condensed structural formula for each of the following:

- propionaldehyde
- 3,4-dimethylhexanal
- 4-bromobutane

---

## 17.6 Carboxylic Acids and Esters

**LEARNING GOAL** Write the IUPAC and common names for carboxylic acids and esters; draw the condensed or line-angle structural formulas when given their names.

In the functional group of a **carboxylic acid**, a carbonyl group is attached to a hydroxyl group \((-OH)\), which forms a **carboxyl group**. Some ways to represent the carboxyl group in propanoic acid follow.

\[
\text{CH}_3-\text{CH}_2-\text{C} \rightarrow \text{OH} \quad \text{Propanoic acid (propionic acid)} \quad \text{CH}_3-\text{CH}_2-\text{COOH}
\]

### IUPAC Names of Carboxylic Acids

The IUPAC names of carboxylic acids replace the \( e \) in the corresponding alkane name with \( oic \) acid. If there are substituents, the carbon chain is numbered beginning with the carboxyl carbon.

\[
\begin{align*}
\text{CH}_3-\text{C} \rightarrow \text{OH} & \quad \text{Methanoic acid} \\
\text{CH}_3-\text{CH} \rightarrow \text{C} \rightarrow \text{OH} & \quad \text{2-Methylpropanoic acid} \\
\text{CH}_3-\text{CH}-\text{CH}_2-\text{C} \rightarrow \text{OH} & \quad \text{3-Hydroxybutanoic acid}
\end{align*}
\]

---

The sour taste of vinegar is due to ethanoic acid (acetic acid).
As with the aldehydes, carboxylic acids with one to four carbon atoms have common names, which are derived from their natural sources. The common names use the prefixes of form, acet, propion, and butyr; these common names are derived from Latin or Greek words (see Table 17.7).

### Table 17.7 IUPAC and Common Names of Selected Carboxylic Acids

<table>
<thead>
<tr>
<th>Condensed Structural Formula</th>
<th>Line-Angle Formula</th>
<th>IUPAC Name</th>
<th>Common Name</th>
<th>Ball-and-Stick Model</th>
</tr>
</thead>
<tbody>
<tr>
<td>H—C—OH</td>
<td>H—C—OH</td>
<td>Methanoic acid</td>
<td>Formic acid</td>
<td></td>
</tr>
<tr>
<td>CH₃—C—OH</td>
<td>CH₃—C—OH</td>
<td>Ethanoic acid</td>
<td>Acetic acid</td>
<td></td>
</tr>
<tr>
<td>CH₃—CH₂—C—OH</td>
<td>CH₃—CH₂—C—OH</td>
<td>Propanoic acid</td>
<td>Propionic acid</td>
<td></td>
</tr>
<tr>
<td>CH₃—CH₂—CH₂—C—OH</td>
<td>CH₃—CH₂—CH₂—C—OH</td>
<td>Butanoic acid</td>
<td>Butyric acid</td>
<td></td>
</tr>
</tbody>
</table>

The simplest aromatic carboxylic acid is benzoic acid. The carbon of the carboxyl group is bonded to carbon 1 in the ring, and the ring is numbered to give the lowest numbers for any substituent.

- Benzoic acid
- 4-Aminobenzoic acid
- 3-Chlorobenzoic acid

### Sample Problem 17.12 Naming Carboxylic Acids

Write the IUPAC name for the following:

- \(\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}\)

#### Try It First

**SOLUTION**

<table>
<thead>
<tr>
<th>Analyze the Problem</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>four-carbon chain, methyl substituent</td>
<td>IUPAC name</td>
<td>position of —CH₃, replace e in alkane name with oic acid</td>
<td></td>
</tr>
</tbody>
</table>
Identify the longest carbon chain and replace the e in the alkane name with oic acid. The IUPAC name of a carboxylic acid with four carbons is butanoic acid.

Step 2 Name and number any substituents by counting the carboxyl carbon as 1. Counting from the right, the methyl substituent is on carbon 2. The IUPAC name is 2-methylbutanoic acid.

Write the IUPAC name for the following:

\[
\text{O} \quad \text{C} \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{OH} \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{C} \quad \text{OH}
\]

Answer: pentanoic acid

**CHEMISTRY LINK TO HEALTH**

Carboxylic Acids in Metabolism

Several carboxylic acids are part of the metabolic processes within our cells. For example, during glycolysis, a molecule of glucose is broken down into two molecules of pyruvic acid, or actually, its carboxylate ion, pyruvate. During strenuous exercise when oxygen levels are low (anaerobic), pyruvic acid is reduced to give lactic acid or the lactate ion.

During exercise, pyruvic acid is converted to lactic acid in the muscles.

In the **citric acid cycle**, also called the Krebs cycle, di- and tricarboxylic acids are oxidized and decarboxylated (loss of CO\(_2\)) to produce energy for the cells of the body. These carboxylic acids are normally referred to by their common names. At the start of the citric acid cycle, citric acid with six carbons is converted to five-carbon \(\alpha\)-ketoglutaric acid. Citric acid is also the acid that gives the sour taste to citrus fruits such as lemons and grapefruits.

The citric acid cycle continues as \(\alpha\)-ketoglutaric acid loses CO\(_2\) to give a four-carbon succinic acid. Then a series of reactions converts succinic acid to oxaloacetic acid. We see that some of the functional groups we have studied along with reactions such as hydration and oxidation are part of the metabolic processes that take place in our cells.
Esters

In the functional group of an ester, the \(-\text{H}\) of the carboxylic acid is replaced by an alkyl group. Fats and oils in our diets contain esters of glycerol and fatty acids, which are long-chain carboxylic acids. The pleasant aromas and flavors of many fruits including bananas, oranges, and strawberries are due to esters.

\[
\begin{align*}
\text{Carboxylic Acid} & \quad \text{Ester} \\
\text{Ethanoic acid} \quad \text{(acetic acid)} & \quad \text{Methyl ethanoate} \quad \text{(methyl acetate)}
\end{align*}
\]

Esterification

In a reaction called esterification, an ester is produced when a carboxylic acid and an alcohol react in the presence of an acid catalyst (usually H\(_2\)SO\(_4\)) and heat. An excess of the alcohol reactant is used to shift the equilibrium in the direction of the formation of the ester product. In this esterification reaction, the \(-\text{OH}\) removed from the carboxylic acid and the \(-\text{H}\) removed from the alcohol combine to form water.

\[
\begin{align*}
\text{Ethanoic acid} \quad \text{(acetic acid)} & + \text{Methanol} \quad \text{(methyl alcohol)} & \rightarrow & \text{Methyl ethanoate} \quad \text{(methyl acetate)} \\
\text{H}^+, \text{heat} & & & \\
\end{align*}
\]
SAMPLE PROBLEM 17.13 Writing Esterification Equations

An ester that has the smell of pineapples can be synthesized from butanoic acid and methanol. Write the balanced chemical equation for the formation of this ester.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>butanoic acid,</td>
<td>esterification</td>
<td>ester and H₂O</td>
</tr>
<tr>
<td></td>
<td>methanol</td>
<td>equation</td>
<td></td>
</tr>
</tbody>
</table>

\[
\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{C} \text{OH} + \text{H} - \text{O} - \text{CH}_3 \rightleftharpoons \text{CH}_3\text{CH}_2\text{CH}_2\text{C} \text{O} - \text{CH}_3 + \text{H}_2\text{O}
\]

**STUDY CHECK 17.13**

What are the IUPAC names of the carboxylic acid and alcohol that are needed to form the following ester, which has the odor of apples? (*Hint:* Separate the O and C=O of the ester group and add H= and —OH to give the original alcohol and carboxylic acid.)

\[
\text{CH}_3\text{CH}_2\text{C} \text{O} - \text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3
\]

**ANSWER**

propanoic acid and 1-pentanol

**Naming Esters**

The name of an ester consists of two words that are derived from the names of the alcohol and the acid in that ester. The first word indicates the alkyl part from the alcohol. The second word is the carboxylate part from the carboxylic acid. The IUPAC names of esters use the IUPAC names for the carbon chain of the acid, while the common names of esters use the common names of the acids. Let’s take a look at the following ester, which has a fruity odor. We start by separating the ester bond into two parts, which gives us the alkyl of the alcohol and the carboxylate of the acid. Then we name the ester as an alkyl carboxylate (see **FIGURE 17.9**).

**FIGURE 17.9** The ester methyl ethanoate (methyl acetate) is made from methyl alcohol and ethanoic acid (acetic acid).

What change is made in the name of the carboxylic acid used to make the ester?

<table>
<thead>
<tr>
<th>IUPAC (common)</th>
<th>Ethanoic acid (acetic acid)</th>
<th>+</th>
<th>Methanol (methyl alcohol)</th>
<th>=</th>
<th>Methyl ethanoate (methyl acetate)</th>
</tr>
</thead>
</table>
Esters in Plants

Many of the fragrances of perfumes and flowers and the flavors of fruits are due to esters. Small esters are volatile, so we can smell them, and soluble in water, so we can taste them. Several esters and their flavors and odors are listed in Table 17.8.

### Table 17.8 Some Esters in Fruits and Flavorings

<table>
<thead>
<tr>
<th>Condensed Structural Formula and Name</th>
<th>Flavor/Odor</th>
</tr>
</thead>
<tbody>
<tr>
<td>O CH₃ — C — O — CH₂—CH₂—CH₃</td>
<td>Pears</td>
</tr>
<tr>
<td>Propyl ethanoate (propyl acetate)</td>
<td></td>
</tr>
<tr>
<td>O CH₃ — C — O — CH₂—CH₂—CH₂—CH₂—CH₃</td>
<td>Bananas</td>
</tr>
<tr>
<td>Pentyl ethanoate (pentyl acetate)</td>
<td></td>
</tr>
<tr>
<td>O CH₃ — C — O — CH₂—CH₂—CH₂—CH₂—CH₂—CH₃</td>
<td>Oranges</td>
</tr>
<tr>
<td>Octyl ethanoate (octyl acetate)</td>
<td></td>
</tr>
<tr>
<td>O CH₃—CH₂—CH₂—C — O —CH₂—CH₃</td>
<td>Pineapples</td>
</tr>
<tr>
<td>Ethyl butanoate (ethyl butyrate)</td>
<td></td>
</tr>
<tr>
<td>O CH₃—CH₂—CH₂—C — O —CH₂—CH₂—CH₂—CH₂—CH₃</td>
<td>Apricots</td>
</tr>
<tr>
<td>Pentyl butanoate (pentyl butyrate)</td>
<td></td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 17.14 Naming Esters**

Write the IUPAC and common names for the following:

O CH₃—CH₂—C — O —CH₂—CH₂—CH₂—CH₃

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>three-carbon carboxylate and propyl</td>
<td>IUPAC name</td>
<td>name alkyl, replace ic acid with ate</td>
</tr>
</tbody>
</table>

**STEP 1** Write the name for the carbon chain from the alcohol as an alkyl group.

The alcohol that is used for the ester is propanol, which is named as the alkyl group propyl.

O CH₃—CH₂—C — O —CH₂—CH₂—CH₂—CH₃ propyl

**STEP 2** Change the ic acid of the acid name to ate.
**STEP 2** Change the *ic acid* of the acid name to *ate*. The carboxylic acid with three carbon atoms is propanoic acid. Replacing the *ic acid* with *ate* gives the IUPAC name propyl propanoate. The common name of propionic acid gives the common name for the ester of propyl propionate.

![Propyl propanoate](image)

**STUDY CHECK 17.14**
Write the IUPAC name for the following ester found in grapes:

![Ethyl heptanoate](image)

**ANSWER**
ethyl heptanoate

---

**QUESTIONS AND PROBLEMS**

**17.6 Carboxylic Acids and Esters**

**LEARNING GOAL** Write the IUPAC and common names for carboxylic acids and esters; draw the condensed or line-angle structural formulas when given their names.

**17.41** Write the IUPAC and common name (if any) for each of the following carboxylic acids:

a. CH₃ — C — OH  
   [Butanoic acid]

b. CH₃ — CH₂ — C — OH  
   [Ethanoic acid]

c. CH₃ — CH₂ — CH₂ — COOH  
   [Valerianic acid]

d. C₆H₅ — COOH  
   [Benzoic acid]

**17.42** Write the IUPAC and common name (if any) for each of the following carboxylic acids:

a. H — C — OH  
   [Formic acid]

b. CH₃ — CH₂ — C — OH  
   [Propanoic acid]

c. Br — C — OH  
   [2-Chloropropionic acid]

d. C₆H₅ — COOH  
   [Benzoic acid]

**17.43** Draw the condensed structural formula for each of the following carboxylic acids:

a. Butanoic acid  
   ![Butanoic acid](image)

b. Benzoic acid  
   ![Benzoic acid](image)

c. 2-Chloroethanoic acid  
   ![2-Chloroethanoic acid](image)

d. 3-Hydroxypropanoic acid  
   ![3-Hydroxypropanoic acid](image)

**17.44** Draw the condensed structural formula for each of the following carboxylic acids:

a. 2-Methylhexanoic acid  
   ![2-Methylhexanoic acid](image)

b. 3-Ethylbenzoic acid  
   ![3-Ethylbenzoic acid](image)

c. 2-Hydroxyacetic acid  
   ![2-Hydroxyacetic acid](image)

d. 2,4-Dibromobutanoic acid  
   ![2,4-Dibromobutanoic acid](image)

**17.45** Draw the condensed structural formula for the ester formed when each of the following reacts with ethyl alcohol:

a. Acetic acid  
   ![Acetic acid with ethyl alcohol](image)

b. Butyric acid  
   ![Butyric acid with ethyl alcohol](image)

**17.46** Draw the condensed structural formula for the ester formed when each of the following reacts with methyl alcohol:

a. Formic acid  
   ![Formic acid with methyl alcohol](image)

b. Propionic acid  
   ![Propionic acid with methyl alcohol](image)

**17.47** Draw the structural formula for the ester formed when the following carboxylic acids and alcohols react:

a. ![Ester formation with ethyl alcohol](image)

b. ![Ester formation with methyl alcohol](image)
17.48 Draw the structural formula for the ester formed when the following carboxylic acids and alcohols react:

a. \[
\text{H}_2\text{COOH} + \text{HOCH}_2\text{CH}_2\text{CH}_2\text{CH}_3 \rightarrow \text{H}_2\text{COOCH}_2\text{CH}_2\text{CH}_2\text{CH}_3
\]
b. \[
\text{C}_6\text{H}_5\text{COOH} + \text{HOCH}_2\text{CH}_2\text{CH}_2\text{CH}_3 \rightarrow \text{C}_6\text{H}_5\text{COOCH}_2\text{CH}_2\text{CH}_2\text{CH}_3
\]

17.49 Write the IUPAC and common names, if any, for each of the following:

a. \[
\text{H} - \text{C} - \text{O} - \text{CH}_3
\]
b. \[
\text{H} - \text{C} - \text{O} - \text{CH}_3
\]
c. \[
\text{CH}_3 - \text{CH}_2 - \text{C} - \text{O} - \text{CH}_2 - \text{CH}_3
\]

17.50 Write the IUPAC and common names, if any, for each of the following:

a. \[
\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{C} - \text{O} - \text{CH}_3
\]
b. \[
\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{C} - \text{O} - \text{CH}_3
\]
c. \[
\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{C} - \text{O} - \text{CH}_3
\]

17.51 Draw the condensed structural formula for each of the following:

a. methyl acetate
b. butyl formate
c. ethyl pentanoate
d. propyl propanoate

17.52 Draw the condensed structural formula for each of the following:

a. hexyl acetate
b. ethyl formate
c. ethyl hexanoate
d. methyl benzoate

17.7 Amines and Amides

**LEARNING GOAL** Write the common names for amines and the IUPAC and common names for amides; draw the condensed or line-angle structural formulas when given their names.

In the functional group of a amine, a nitrogen atom is attached to one or more carbon atoms. In methylamine, a methyl group replaces one hydrogen atom in ammonia. The bonding of two methyl groups gives dimethylamine. In trimethylamine, methyl groups replace all three hydrogen atoms attached to the nitrogen atom (see **FIGURE 17.10**).

**Naming Amines**

Several systems are used for naming amines. For simple amines, the common names are often used. In the common name, the alkyl groups bonded to the nitrogen atom are listed in alphabetical order. The prefixes *di* and *tri* are used to indicate two and three identical substituents.

\[
\begin{align*}
\text{H} & \quad \text{N} & \quad \text{H} \\
\text{CH}_3 & \quad \text{N} & \quad \text{H} \\
\text{CH}_3 & \quad \text{N} & \quad \text{CH}_3 \\
\text{CH}_3 & \quad \text{N} & \quad \text{CH}_3
\end{align*}
\]

**FIGURE 17.10** Amines have one or more carbon atoms bonded to the N atom.

Q How many carbon atoms are bonded to the nitrogen atom in dimethylamine?
Aromatic Amines

The aromatic amines use the name *aniline*, which is approved by IUPAC.

Aniline is used to make many dyes, which give color to wool, cotton, and silk fibers, as well as blue jeans. It is also used to make the polymer polyurethane and in the synthesis of the pain reliever acetaminophen.

Indigo used in blue dyes can be obtained from tropical plants such as *Indigofera tinctoria*.

**SAMPLE PROBLEM 17.15 Naming Amines**

Write the common name for each of the following amines:

\[ \text{CH}_3 \]

a. \( \text{CH}_3 - \text{CH}_2 - \text{NH}_2 \)        b. \( \text{CH}_3 - \text{N} - \text{CH}_3 \)

**TRY IT FIRST**

**SOLUTION**

a. This amine has one ethyl group attached to the nitrogen atom; its name is ethylamine.

b. The common name for an amine with three methyl groups attached to the nitrogen atom is trimethylamine.

**STUDY CHECK 17.15**

What is the name of the following amine?

\[ \text{NH}_2 \]

**ANSWER**
aniline
Alkaloids are physiologically active nitrogen-containing compounds produced by plants. The term *alkaloid* refers to the “alkali-like” or basic characteristics of amines. Certain alkaloids are used in anesthetics, in antidepressants, and as stimulants, and many are habit forming.

As a stimulant, nicotine increases the level of adrenaline in the blood, which increases the heart rate and blood pressure. Nicotine is addictive because it activates pleasure centers in the brain. Coniine, which is obtained from hemlock, is extremely toxic.

Nicotine

\[
\text{N} \quad \text{H} \quad \text{CH} \quad \text{H}
\]

Coniine

Caffeine is a central nervous system stimulant. Present in coffee, tea, soft drinks, energy drinks, chocolate, and cocoa, caffeine increases alertness, but it may cause nervousness and insomnia. Caffeine is also used in certain pain relievers to counteract the drowsiness caused by an antihistamine.

Caffeine

\[
\text{O} \quad \text{N} \quad \text{C} \quad \text{CH} \quad \text{CH}
\]

Several alkaloids are used in medicine. Quinine, obtained from the bark of the cinchona tree, has been used in the treatment of malaria since the 1600s. Atropine from nightshade (belladonna) is used in low concentrations to accelerate slow heart rates and as an anesthetic for eye examinations.

Quinine

\[
\text{O} \quad \text{C} \quad \text{CH} \quad \text{H}
\]

Atropine

For many centuries morphine and codeine, alkaloids found in the oriental poppy plant, have been used as effective painkillers. Codeine, which is structurally similar to morphine, is used in some prescription painkillers and cough syrups. Heroin, obtained by a chemical modification of morphine, is strongly addicting and is not used medically. The structure of the prescription drug OxyContin (oxycodone) used to relieve severe pain is similar to heroin. Today, there are an increasing number of deaths from OxyContin abuse because its physiological effects are also similar to those of heroin.

Morphine

\[
\text{CH} \quad \text{C} \quad \text{O} \quad \text{N} \quad \text{CH}
\]

Codeine

\[
\text{CH} \quad \text{C} \quad \text{O} \quad \text{N} \quad \text{CH}
\]

Heroin

\[
\text{CH} \quad \text{C} \quad \text{O} \quad \text{N} \quad \text{CH}
\]

OxyContin

The green, unripe poppy seed capsule contains a milky sap (opium) that is the source of the alkaloids morphine and codeine.
Amides

In the functional group of an amide, the hydroxyl group of a carboxylic acid is replaced by a nitrogen atom (see FIGURE 17.11).

Amidation

An amide is produced in a reaction called amidation, in which a carboxylic acid reacts with ammonia or an amine. A molecule of water is eliminated, and the fragments of the carboxylic acid and amine molecules join to form the amide, much like the formation of esters.

SAMPLE PROBLEM 17.16 Formation of Amides

Draw the condensed structural formula for the amide product in the following reaction:

\[
\text{CH}_3\text{C} = \text{OH} + \text{H}_2\text{N} \text{CH}_2\text{CH}_3 \xrightarrow{\text{Heat}} \text{CH}_3\text{C} = \text{N} \text{CH}_2\text{CH}_3
\]

SOLUTION

The condensed structural formula for the amide product can be drawn by attaching the carbonyl group from the acid to the nitrogen atom of the amine. The \(-\text{OH}\) group is removed from the acid and \(-\text{H}\) from the amine to form water.

\[
\text{CH}_3\text{C} = \text{N} \text{CH}_2\text{CH}_3
\]
STUDY CHECK 17.16
Draw the condensed structural formulas for the carboxylic acid and amine needed to prepare the following amide:

\[
\begin{align*}
&\text{O} \quad \text{CH}_3 \\
&\text{H} - \text{C} - \text{N} - \text{CH}_3 \\
\end{align*}
\]

ANSWER

\[
\begin{align*}
&\text{O} \quad \text{CH}_3 \\
&\text{H} - \text{C} - \text{OH} \quad \text{and} \quad \text{H} - \text{N} - \text{CH}_3 \\
\end{align*}
\]

Naming Amides

In the IUPAC and common names for amides, the \textit{oic acid} or \textit{ic acid} from the carboxylic acid name is replaced with \textit{amide}. We can diagram the name of an amide in the following way:

From butanoic acid (butyric acid)

\[
\begin{align*}
\text{CH}_3 \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{C} \quad \text{NH}_2 \\
\text{IUPAC (Common)} \quad \text{Butanamide (butyramide)} \\
\end{align*}
\]

From ammonia

\[
\begin{align*}
\text{O} \\
\text{H} - \text{C} - \text{NH}_2 \\
\text{Methanamide (formamide)} \quad \text{CH}_3 \quad \text{C} \quad \text{NH}_2 \\
\text{Ethanamide (acetamide)} \quad \text{O} \quad \text{C} \quad \text{NH}_2 \\
\text{Benzamide} \\
\end{align*}
\]

SAMPLE PROBLEM 17.17 Naming Amides

Write the IUPAC and common names for

\[
\begin{align*}
&\text{O} \\
&\text{CH}_3 - \text{CH}_2 - \text{C} - \text{NH}_2 \\
\end{align*}
\]

TRY IT FIRST

SOLUTION

The IUPAC name of the carboxylic acid is propanoic acid; the common name is propionic acid. Replacing the \textit{oic acid} or \textit{ic acid} ending with \textit{amide} gives the IUPAC name of propanamide and common name of propionamide.

STUDY CHECK 17.17

Write the IUPAC name for

\[
\begin{align*}
&\text{O} \\
&\text{NH}_2 \\
\end{align*}
\]

ANSWER

hexanamide
LEARNING GOAL Write the common names for amines and the IUPAC and common names for amides; draw the condensed or line-angle structural formulas when given their names.

17.53 Write the common name for each of the following:
   a. CH₃—NH₂
   b. CH₃—NH—CH₂—CH₃
   c. CH₃—CH₂—N—CH₂—CH₃

17.54 Write the common name for each of the following:
   a. CH₃—CH₂—NH₂
   b. CH₃—N—CH₂—CH₃
   c. NH₂

17.55 Draw the condensed structural formula for the amide formed in each of the following reactions:
   a. CH₃—C—OH + NH₃ \( \xrightarrow{\text{Heat}} \)
   b. CH₃—C—OH + H₂N—CH₂—CH₃ \( \xrightarrow{\text{Heat}} \)
   c. CH₃—C—OH + H₂N—CH₂—CH₂—CH₃ \( \xrightarrow{\text{Heat}} \)

17.56 Draw the condensed structural formula for the amide formed in each of the following reactions:
   a. CH₃—CH₂—CH₂—CH₂—C—OH + NH₃ \( \xrightarrow{\text{Heat}} \)
   b. CH₃—CH—C—OH + H₂N—CH₂—CH₃ \( \xrightarrow{\text{Heat}} \)
   c. CH₃—CH₂—C—OH + H₂N — \( \xrightarrow{\text{Heat}} \)

17.57 Write the IUPAC and common name (if any) for each of the following amides:
   a. CH₃—C—NH₂
   b. CH₃—CH₂—CH—C—NH₂
   c. H—C—NH₂

17.58 Write the IUPAC and common name (if any) for each of the following amides:
   a. CH₃—CH—C—NH₂
   b. CH₃—CH₂—CH₂—CH₂—C—NH₂
   c. CH₃—O

17.59 Draw the condensed structural formula for each of the following amides:
   a. propionamide
   b. 2-methylpentanamide
   c. methanamide

17.60 Draw the condensed structural formula for each of the following amides:
   a. heptanamide
   b. benzamide
   c. 3-methylbutanamide
Follow Up

DIANE’S TREATMENT IN THE BURN UNIT

When Diane arrived at the hospital, she was transferred to the ICU burn unit. She had second-degree burns that caused damage to the underlying layers of skin and third-degree burns that caused damage to all the layers of the skin. Because body fluids are lost when deep burns occur, a lactated Ringer’s solution was administered. The most common complications of burns are related to infection. To prevent infection, her skin was covered with a topical antibiotic. The next day Diane was placed in a tank to remove dressings, lotions, and damaged tissue. Dressings and ointments were changed every eight hours. Over a period of 3 months, grafts of Diane’s unburned skin were used to cover burned areas. She remained in the burn unit for 3 months and then was discharged. However, Diane returned to the hospital for more skin grafts and plastic surgery.

The arson investigators determined that gasoline was the primary accelerant used to start the fire at Diane’s house. Because there was a lot of paper and dry wood in the area, the fire spread quickly. Some of the hydrocarbons found in gasoline include hexane, heptane, octane, nonane, decane, and toluene.

Applications

17.61 The medications for Diane were contained in blister packs made of polychlorotrifluoroethylene (PCTFE). Draw the expanded structural formula for a portion of PCTFE polymer from three monomers of chlorotrifluoroethylene shown below:

![Chlorotrifluoroethylene](image)

Tablets are contained in blister packs made from polychlorotrifluoroethylene (PCTFE).

17.62 New polymers have been synthesized to replace PVC in IV bags and cling film. One of these is EVA, ethylene polyvinylacetate. Draw the expanded structural formula for a portion of EVA using an alternating sequence that has two of each monomer shown below:

![Ethylene and Vinyl acetate](image)

17.63 Write the balanced chemical equation for the complete combustion of each of the following hydrocarbons found in gasoline used to start a fire:

a. nonane
b. toluene

17.64 Write the balanced chemical equation for the complete combustion of each of the following hydrocarbons found in gasoline used to start a fire:

a. hexane
b. octane
17.1 Alkanes

LEARNING GOAL Write the IUPAC names and draw the condensed or line-angle structural formulas for alkanes.

- Alkanes are hydrocarbons that have only C—C single bonds.
- In the expanded structural formula, a separate line is drawn for every bonded atom.
- A condensed structural formula depicts groups composed of each carbon atom and its attached hydrogen atoms.
- A line-angle structural formula represents the carbon skeleton as ends and corners of a zigzag line.
- Substituents such as alkyl groups and halogen atoms (named as fluoro, chloro, bromo, or iodo) can replace hydrogen atoms on the main chain.
- The IUPAC system is used to name organic compounds by indicating the number of carbon atoms and the position of any substituents.

17.2 Alkenes, Alkynes, and Polymers

LEARNING GOAL Write the IUPAC names and draw the condensed or line-angle structural formulas for alkenes and alkynes.

- Alkenes are unsaturated hydrocarbons that contain carbon–carbon double bonds (C=C).
- Alkynes contain a carbon–carbon triple bond (C≡C).
- The IUPAC names of alkenes end with ene; alkyne names end with yne.
- The main chain is numbered from the end nearer the double or triple bond.
- Alkenes react with hydrogen using a metal catalyst to form alkanes.
- Polymers are long-chain molecules that consist of many repeating units of smaller carbon molecules called monomers.

17.3 Aromatic Compounds

LEARNING GOAL Describe the bonding in benzene; name aromatic compounds, and draw their line-angle structural formulas.

- Most aromatic compounds contain benzene, C₆H₆, a cyclic structure containing six carbon atoms and six hydrogen atoms.
- The structure of benzene is represented as a hexagon with a circle in the center.
- The IUPAC system uses the names of benzene, toluene, aniline, and phenol.
- In the IUPAC name, two or more substituents are numbered and listed in alphabetical order.
17.4 Alcohols and Ethers

**LEARNING GOAL** Write the IUPAC and common names for alcohols and ethers; draw the condensed or line-angle structural formulas when given their names.

- The functional group of an alcohol is the hydroxyl group (―OH) bonded to a carbon chain.
- In a phenol, the hydroxyl group is bonded to an aromatic ring.
- In the IUPAC system, the names of alcohols have "ol" endings, and the location of the ―OH group is given by numbering the carbon chain.
- In an ether, an oxygen atom (―O―) is connected by single bonds to two alkyl or aromatic groups.
- In the common names of ethers, the alkyl groups are listed alphabetically followed by the name "ether." 

17.5 Aldehydes and Ketones

**LEARNING GOAL** Write the IUPAC and common names for aldehydes and ketones; draw the condensed or line-angle structural formulas when given their names.

- Aldehydes and ketones contain a carbonyl group (C=O), which is strongly polar.
- In aldehydes, the carbonyl group appears at the end of the carbon chain attached to at least one hydrogen atom.
- In ketones, the carbonyl group occurs between two carbon atoms.
- In the IUPAC system, the "e" in the alkane name is replaced with "al" for aldehydes and "one" for ketones.
- For ketones with more than four carbon atoms in the main chain, the carbonyl group is numbered to show its location.
- Many of the simple aldehydes and ketones use common names.

17.6 Carboxylic Acids and Esters

**LEARNING GOAL** Write the IUPAC and common names for carboxylic acids and esters; draw the condensed or line-angle structural formulas when given their names.

- A carboxylic acid contains the carbonyl functional group, which is a hydroxyl group connected to a carbonyl group.
- The IUPAC name of a carboxylic acid is obtained by replacing the "e" in the alkane name with "oic acid."
- The common names of carboxylic acids with one to four carbon atoms are: formic acid, acetic acid, propionic acid, and butyric acid.
- In an ester, a carbon atom replaces the H of the hydroxyl group of a carboxylic acid.
- When heated in the presence of a strong acid, a carboxylic acid reacts with an alcohol to produce an ester. A molecule of water is removed: ―OH from the carboxylic acid and ―H from the alcohol molecule.
- The names of esters consist of two words, the alkyl group from the alcohol and the name of the carboxylate obtained by replacing the "ic acid" ending with "ate."

17.7 Amines and Amides

**LEARNING GOAL** Write the common names for amines and the IUPAC and common names for amides; draw the condensed or line-angle structural formulas when given their names.

- A nitrogen atom attached to one or more carbon atoms forms an amine.
- In the common names of simple amines, the alkyl groups are listed alphabetically followed by the suffix "amine."
- Amides are derivatives of carboxylic acids in which the hydroxyl group is replaced by a nitrogen atom. Amides are named by replacing the "ic acid" or "oic acid" ending with "amide."
- Amides are formed by the reaction of a carboxylic acid with ammonia or an amine.

### SUMMARY OF NAMING

<table>
<thead>
<tr>
<th>Condensed Structural Formula</th>
<th>Family</th>
<th>IUPAC Name</th>
<th>Common Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₃—CH₂—CH₃</td>
<td>Alkane</td>
<td>Propane</td>
<td></td>
</tr>
<tr>
<td>CH₃</td>
<td>Alkane with a substituent</td>
<td>Methylpropane</td>
<td></td>
</tr>
<tr>
<td>CH₃—CH=CH₂</td>
<td>Alkene</td>
<td>Propene</td>
<td>Propylene</td>
</tr>
<tr>
<td>CH₃—C≡CH</td>
<td>Alkyne</td>
<td>Propyne</td>
<td></td>
</tr>
<tr>
<td>[Aromatic structure]</td>
<td>Aromatic</td>
<td>Benzene</td>
<td></td>
</tr>
</tbody>
</table>
### SUMMARY OF REACTIONS

The chapter sections to review are shown after the name of the reaction.

#### Combustion (17.1)

\[
\text{CH}_3\text{CH}_2\text{CH}_3(g) + 5\text{O}_2(g) \xrightarrow{\Delta} 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g) + \text{energy}
\]

**Propane**

#### Hydrogenation (17.2)

\[
\text{H}_2\text{C} = \text{CH} - \text{CH}_3 + \text{H}_2 \xrightarrow{\text{Pt}} \text{CH}_3 - \text{CH}_{2} - \text{CH}_3
\]

**Propene**

**Propane**

#### Esterification (17.2)

\[
\text{CH}_3 - \text{C} - \text{OH} + \text{HO} - \text{CH}_3 \xrightarrow{\text{H}^+; \text{heat}} \text{CH}_3 - \text{C} - \text{O} - \text{CH}_3 + \text{H}_2\text{O}
\]

**Ethanoic acid** (acetic acid)

**Methanol** (methyl alcohol)

**Methyl ethanoate** (methyl acetate)

#### Amidation (17.7)

\[
\text{CH}_3 - \text{CH}_2 - \text{C} - \text{OH} + \text{H} - \text{N} - \text{H} \xrightarrow{\text{Heat}} \text{CH}_3 - \text{CH}_2 - \text{C} - \text{N} - \text{H} + \text{H}_2\text{O}
\]

**Propanoic acid** (propionic acid)

**Ammonia**

**Propanamide** (propionamide)

### KEY TERMS

**alcohol** An organic compound that contains a hydroxyl group (—OH) attached to a carbon chain.

**aldehyde** An organic compound that contains a carbonyl group (C=O) bonded to at least one hydrogen atom.

**alkane** A type of hydrocarbon in which the carbon atoms are connected only by single bonds.

**alkene** A type of hydrocarbon that contains a carbon-carbon double bond (C=C).

**alkyl group** An alkane minus one hydrogen atom. Alkyl groups are named like the corresponding alkanes except a *yl* ending replaces *ane.*
alkyne A type of hydrocarbon that contains a carbon–carbon triple bond (C≡C).

amide An organic compound in which the hydroxyl group of a carboxylic acid is replaced by a nitrogen atom.

amine An organic compound that contains a nitrogen atom bonded to one or more carbon atoms.

aromatic compound A compound that contains the ring structure of benzene.

benzene A ring of six carbon atoms each of which is attached to one hydrogen atom, C₆H₆.

carboxylic acid An organic compound that contains the carboxyl group (—COOH).

condensed structural formula A formula that shows the carbon atoms grouped with the attached number of hydrogen atoms.

ester An organic compound in which the —H of a carboxyl group is replaced by a carbon atom.

ether An organic compound in which an oxygen atom is bonded to two carbon atoms.

expanded structural formula A formula that shows all of the atoms and the bonds connected to each atom.

**CORE CHEMISTRY SKILLS**

The chapter section containing each Core Chemistry Skill is shown in parentheses at the end of each heading.

**Naming and Drawing Alkanes (17.1)**

- The alkanes ethane, propane, and butane contain two, three, and four carbon atoms, respectively, connected in a row or a continuous chain.
- Alkanes with five or more carbon atoms in a chain are named using the prefixes pent (5), hex (6), hept (7), oct (8), non (9), and dec (10).
- In the condensed structural formula, the carbon and hydrogen atoms on the ends are written as —CH₃ and the carbon and hydrogen atoms in the middle are written as —CH₂—.

**Example:**

**a.** What is the name of CH₃—CH₂—CH₂—CH₃?

**Answer:** a. CH₃—C≡CH—CH₃

**b.** When H₂ adds to a double bond, the product is an alkane with the same number of carbon atoms.

**Answer:** CH₃—CH—CH₂—CH₃

**Naming Aldehydes and Ketones (17.5)**

- In the IUPAC system, an aldehyde is named by replacing the e in the alkane name with al and a ketone by replacing the e with one.
- The position of a substituent on an aldehyde is indicated by numbering the carbon chain from the carbonyl group and given in front of the name.
- For a ketone, the carbon chain is numbered from the end nearer the carbonyl group.

**Example:** Give the IUPAC name for the following:

**Answer:** 5-methyl-3-hexanone

**Writing Equations for Hydrogenation and Polymerization (17.2)**

- The addition of small molecules to the double bond is a characteristic reaction of alkenes.
- Hydrogenation adds hydrogen atoms to the double bond of an alkene, using a metal catalyst, to form an alkane.
- In polymerization, many small molecules join together to form a polymer.

**Example:**

**a.** Draw the condensed structural formula for 2-methyl-2-butene.

**b.** Draw the condensed structural formula for the product of the hydrogenation of 2-methyl-2-butene.

**Naming Carboxylic Acids (17.6)**

- The IUPAC name of a carboxylic acid is obtained by replacing the e in the corresponding alkane name with oic acid.
- The common names of carboxylic acids with one to four carbon atoms are formic acid, acetic acid, propionic acid, and butyric acid.

**Example:** Write the IUPAC name for the following:

**Answer:** 3-bromobutanoic acid
Forming Esters (17.6)
• Esters are formed when carboxylic acids react with alcohols in the presence of a strong acid and heat.

Example: Draw the condensed structural formula for the ester product of the reaction of propanoic acid and methanol.

Answer: \( \text{CH}_3 \text{CH}_2 \text{O} \text{C} \text{CH}_3 \)

Forming Amides (17.7)
• Amides are formed when carboxylic acids react with ammonia or primary or secondary amines in the presence of heat.

Example: Draw the amide produced by the reaction of propanoic acid and ethylamine.

Answer: \( \text{CH}_3 \text{CH}_2 \text{C} \text{N} \text{CH}_2 \text{CH}_3 \)

UNDERSTANDING THE CONCEPTS

The chapter sections to review are shown in parentheses at the end of each question.

17.65 The polymer Teflon is made from the monomer 1,1,2,2-tetrafluoroethane. Draw the chemical equation of the formation of Teflon from the monomer. (17.2)

17.66 A garden hose is made of polyvinyl chloride (PVC), a polymer of chloroethene (vinyl chloride). Draw a portion of the polymer (use three monomers) for PVC. (17.2)

17.67 Citronellal, found in oil of citronella, lemon, and lemon grass, is used in perfumes and as an insect repellent. (17.2, 17.5)

Citronellal has the following structure:

\( \text{CH}_3 \text{CH} \equiv \text{CH} \text{CH}_2 \text{CH} \equiv \text{CH} \text{CH}_2 \text{O} \text{C} \text{CH}_3 \)

a. Complete the IUPAC name for citronellal: ___-di _____-___-octenal.
b. What does the \( \text{en} \) in octenal signify?
c. What does the \( \text{al} \) in octenal signify?

17.68 Draw the condensed structural formula for each of the following: (17.5)

a. 2-heptanone, an alarm pheromone of bees
b. 2,6-dimethyl-3-heptanone, a communication pheromone of bees

Bees emit chemicals called pheromones to communicate.
17.69 Draw the condensed structural formula for each of the following: (17.1, 17.2)
   a. 3-ethylhexane
   b. 2-pentene
   c. 2-hexyne

17.70 Draw the condensed structural formula for each of the following: (17.1, 17.2)
   a. 1-chloro-2-butyne
   b. 2,3-dimethylpentane
   c. 3-hexene

17.71 Write the IUPAC name for each of the following: (17.1, 17.2)
   a. \[CH_3CH_2CCl_3\]
   b. \[CH_3CH_2CH_2CH_2CH_3\]
   c. \[CH_3CH=CHCH=CHCH_3\]

17.72 Write the IUPAC name for each of the following: (17.1, 17.2)
   a. \[H_2C\equivC\equivCH_2\equivCH_2\equivCH_3\]
   b. \[CH_3CH_2Cl\]
   c. \[CH_3CH_2CH=CHCH=CHCH_3\]

17.73 Classify each of the following according to its functional group: (17.2)
   a. \[\text{O} \vdash \text{C} \equiv \text{O}\]
   b. \[\text{CH}=\text{CH} \equiv \text{C} \equiv \text{H}\]
   c. \[\text{CH}=\text{C} \equiv \text{O} \equiv \text{CH}\]
   d. \[\text{CH}_3\text{CH}=\text{CH} \equiv \text{O} \equiv \text{CH}_3\]

17.74 Classify each of the following according to its functional group: (17.2)
   a. \[\text{CH}_3\text{CH}=\text{CH} \equiv \text{C} \equiv \text{CH}_3\]
   b. \[\text{CH}_3\text{CH}=\text{CH} \equiv \text{CH}_3\]
   c. \[\text{CH}_3\equiv\text{C} \equiv \text{C} \equiv \text{OH}\]
   d. \[\text{CH}_3\equiv\text{C} \equiv \text{C} \equiv \text{H}\]

17.75 Select the class of organic compound that matches each of the following descriptions (a to d): alkane, alkene, alkyne, alcohol, ether, aldehyde, ketone, carboxylic acid, ester, amine, or amide. (17.2)
   a. contains a carbonyl group between two carbons
   b. is a weak acid
   c. contains a carbon–carbon triple bond
   d. produced by a hydrogenation reaction using a metal catalyst with alkene

17.76 Select the class of organic compound that matches each of the following descriptions (a to d): alkane, alkene, alkyne, alcohol, ether, aldehyde, ketone, carboxylic acid, ester, amine, or amide. (17.2)
   a. contains a carbonyl group between two carbons
   b. is a weak acid
   c. contains a carbon–carbon triple bond
   d. produced by a hydrogenation reaction using a metal catalyst with alkene

17.77 Classify each of the following according to its functional group(s): (17.2)
   a. \[\text{C} \equiv \text{H}\]
   b. \[\text{CH}=\text{CH} \equiv \text{C} \equiv \text{H}\]

17.78 Classify each of the following according to its functional group(s): (17.2)
   a. BHA is an antioxidant used as a preservative in foods such as baked goods, butter, meats, and snack foods.
   b. \[\text{CH}_3\text{CH}_3\text{C} \equiv \text{C} \equiv \text{CH}_3\]
   b. \[\text{CH}_3\text{CH}_3\text{C} \equiv \text{C} \equiv \text{CH}_3\]
   c. \[\text{CH}_3\text{CH}_3\text{C} \equiv \text{C} \equiv \text{CH}_3\]

17.79 Name each of the following aromatic compounds: (17.8)
   a. \[\text{CH}_3\text{Cl}\]
   b. \[\text{NO}_2\text{OH}\]
   c. \[\text{CH}_3\text{CH}_3\text{C} \equiv \text{C} \equiv \text{CH}_3\]
17.80 Name each of the following aromatic compounds: (17.8)

a. 

\[
\begin{array}{c}
H_2C \quad \text{Cl} \\
\text{CH}_3
\end{array}
\]

b. 

\[
\begin{array}{c}
H_2C \quad \text{Cl} \\
\text{CH}_3
\end{array}
\]

c. 

\[
\begin{array}{c}
H_2C \quad \text{NH}_2 \\
\text{CH}_3
\end{array}
\]

17.81 Draw the structural formula for each of the following: (17.3)

a. 2-iodotoluene
b. 1,4-dichlorobenzene
c. 1,3,5-tribromobenzene

17.82 Draw the structural formula for each of the following:

(17.3)

a. 2,5-dimethylaniline
b. 4-chloroaniline
c. propanamide

17.83 Write the IUPAC name for a and b and the common name for c. (17.4)

a. 

\[
\begin{array}{c}
\text{CH}_3 \\
\text{CH}_2
\end{array}
\]

b. 

\[
\begin{array}{c}
\text{OH} \\
\text{CH}_3
\end{array}
\]

c. 

\[
\begin{array}{c}
\text{CH}_3 \\
\text{CH}_2 \quad \text{O} \\
\text{CH}_2 \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{CH}_3
\end{array}
\]

17.84 Write the IUPAC name for a and b and the common name for c. (17.4)

a. 

\[
\begin{array}{c}
\text{Cl} \\
\text{OH} \\
\text{CH}_3
\end{array}
\]

b. 

\[
\begin{array}{c}
\text{CH}_3 \\
\text{CH} \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{OH}
\end{array}
\]

c. 

\[
\begin{array}{c}
\text{CH}_3 \\
\text{CH} \quad \text{CH}_2 \\
\text{O}
\end{array}
\]

17.85 Draw the condensed structural formula for each of the following: (17.4)

a. 3-bromo-2-pentanol
b. 2-methyl-1-propanol
c. methyl propyl ether

17.86 Draw the condensed structural formula for each of the following: (17.4)

a. 2-methylphenol
b. 1,1-dichloro-2-hexanol
c. 2-methyl-2-propanol

17.87 Write the IUPAC name for each of the following: (17.5)

a. 

\[
\begin{array}{c}
\text{CH}_3 \\
\text{O} \\
\text{C} \quad \text{H}
\end{array}
\]

b. 

\[
\begin{array}{c}
\text{Cl} \quad \text{H}_2 \quad \text{C} \\
\text{Cl} \quad \text{O}
\end{array}
\]

c. 

\[
\begin{array}{c}
\text{CH}_3 \quad \text{H} \\
\text{C} \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{CH}_3
\end{array}
\]

17.88 Write the IUPAC name for each of the following: (17.5)

a. 

\[
\begin{array}{c}
\text{CH}_3 \quad \text{CH}_2 \quad \text{C} \\
\text{CH}_3
\end{array}
\]

b. 

\[
\begin{array}{c}
\text{O} \\
\text{C} \quad \text{H}
\end{array}
\]

c. 

\[
\begin{array}{c}
\text{CH}_3 \quad \text{CH} \\
\text{C} \quad \text{CH}_2 \quad \text{CH}_3
\end{array}
\]

17.89 Draw the condensed structural formula for each of the following: (17.5)

a. 2-bromobenzaldehyde
b. 3-chloropropionaldehyde
c. ethyl methyl ketone

17.90 Draw the condensed structural formula for each of the following: (17.5)

a. butyraldehyde
b. 2-bromobutanal
c. 3,5-dimethylhexanal

17.91 Write the IUPAC name for each of the following: (17.6)

a. 

\[
\begin{array}{c}
\text{CH}_3 \\
\text{O} \\
\text{C} \quad \text{H}_2 \quad \text{CH}_3
\end{array}
\]

b. 

\[
\begin{array}{c}
\text{O} \\
\text{C} \quad \text{OH}
\end{array}
\]

c. 

\[
\begin{array}{c}
\text{O} \\
\text{CH}_2 \quad \text{CH}_3
\end{array}
\]

17.92 Write the IUPAC name for each of the following: (17.6)

a. 

\[
\begin{array}{c}
\text{CH}_3 \quad \text{CH} \\
\text{C} \quad \text{OH}
\end{array}
\]

b. 

\[
\begin{array}{c}
\text{O} \\
\text{C} \quad \text{OH}
\end{array}
\]

c. 

\[
\begin{array}{c}
\text{O} \\
\text{C} \quad \text{CH}_2 \quad \text{CH}_3
\end{array}
\]
17.93 Draw the condensed structural formula for each of the following: (17.7)
   a. ethylamine
   b. hexanamide
   c. triethylamine
17.94 Draw the condensed structural formula for each of the following: (17.7)
   a. formamide
   b. ethylpropylamine
   c. diethylmethylamine
17.95 Write the name for each of the following: (17.7)
   a. CH₃–CH₂–CH₂–CH₂–C–NH₂
   b. CH₃–CH₂–CH₂–CH₂–NH₂
17.96 Write the name for each of the following: (17.7)
   a. Cl–CH₂–CH₂–CH₂–C–NH₂
   b. CH₃–CH₂–C–NH₂
   c. CH₃–CH₂–CH₂–N–CH₂–CH₃

**CHALLENGE QUESTIONS**

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

17.97 Draw the condensed structural formulas and write the IUPAC names for all eight alcohols that have the molecular formula C₅H₁₂O. (17.4)
17.98 There are four amine isomers with the molecular formula C₃H₇N. Draw their condensed structural formulas. (17.7)
17.99 The insect repellent DEET can be made from an amidation reaction of 3-methylbenzoic acid and diethylamine. Draw the condensed structural formula for DEET. (17.7)

Many insect repellents contain DEET, which is an amide.

17.100 One of the compounds that give blackberries their odor and flavor can be made by heating hexanoic acid and 1-propanol with an acid catalyst. Draw the condensed structural formula for this compound. (17.6)
17.101 Complete and balance each of the following reactions: (17.1, 17.2)
   a. CH₃–CH₂–C≡CH + O₂ →
   b. CH₃–CH₂–CH≡CH + CH₂–CH₂ + H₂ →
17.102 Complete and balance each of the following reactions: (17.1, 17.2)
   a. H₂C≡CH + H₂ →
   b. CH₃–CH₂–CH₂–CH₂–CH≡CH₂ + O₂ →

**ANSWERS**

Answers to Selected Questions and Problems

17.1 a. hexane   b. heptane   c. pentane
17.3 a. same molecule
   b. isomers of C₆H₁₄
   c. isomers of C₆H₁₈
17.5 a. 2-fluorobutane
   b. dimethylpropane
   c. 2-chloro-3-methylpentane
17.9  a. \( \text{CH}_4(g) + 2\text{O}_2(g) \overset{\Delta}{\longrightarrow} \text{CO}_2(g) + 2\text{H}_2\text{O}(g) + \text{energy} \)
   b. \( 2\text{C}_6\text{H}_4(l) + 19\text{O}_2(g) \overset{\Delta}{\longrightarrow} 12\text{CO}_2(g) + 14\text{H}_2\text{O}(g) + \text{energy} \)

17.11  a. An alkene has a double bond.
   b. An alkyne has a triple bond.
   c. An alkene has a double bond.

17.13  a. ethene
   b. 2-methylpropene
   c. 2-pentyne

17.15  a. \( \text{CH}_3 - \text{CH} = \text{CH}_2 \)
   b. \( \text{HC} \equiv \text{C} - \text{CH}_2 - \text{CH}_2 - \text{CH}_3 \)
   c. \( \text{H}_2\text{C} \equiv \text{C} - \text{CH}_2 - \text{CH}_3 \)

17.17  a. \( \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{CH}_3 \)
   b. \( \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_3 \)
   c. \( \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_3 \)

17.19  A polymer is a large molecule composed of small units that are repeated many times.

17.21  \( 3\text{H}_2\text{C} \equiv \text{CH} \rightarrow \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{F} - \text{F} - \text{F} - \text{F} - \text{F} - \text{F} - \text{F} - \text{F} \)

17.23  \( \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{F} - \text{F} - \text{F} - \text{F} - \text{F} - \text{F} - \text{F} - \text{F} \)

17.25  a. 2-chlorotoluene
   b. ethylbenzene
   c. phenol

17.27  a. \( \text{NH}_2 \)
   b. \( \text{Cl} \)
   c. \( \text{CH}_3 \)

17.29  a. ethanol
   b. 2-butanol
   c. 2-chlorophenol

17.31  a. \( \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{OH} \)
   b. \( \text{CH}_3 - \text{OH} \)
   c. \( \text{CH}_3 - \text{CH}_2 - \text{CH} - \text{CH}_2 - \text{CH}_3 \)
   d. \( \text{CH}_3 - \text{C} - \text{CH}_2 - \text{CH}_3 \)

17.33  a. ethyl methyl ether
   b. dipropyl ether
   c. butyl propyl ether

17.35  a. acetaldehyde
   b. methyl propyl ketone
   c. formaldehyde

17.37  a. propanal
   b. 2-methyl-3-pentanone
   c. benzaldehyde

17.39  a. \( \text{CH}_3 - \text{C} - \text{H} \)
   b. \( \text{CH}_3 - \text{C} - \text{CH}_2 - \text{CH}_2 - \text{CH}_3 \)
   c. \( \text{CH}_3 - \text{C} - \text{CH}_2 - \text{CH}_2 - \text{CH}_3 \)

17.41  a. ethanoic acid (acetic acid)
   b. propanoic acid (propionic acid)
   c. 3-methylhexanoic acid
   d. 3-bromobenzoic acid

17.43  a. \( \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{C} - \text{OH} \)
   b. \( \text{C} - \text{OH} \)
   c. \( \text{Cl} - \text{CH}_2 - \text{C} - \text{OH} \)
   d. \( \text{HO} - \text{CH}_2 - \text{CH}_2 - \text{C} - \text{OH} \)

17.45  a. \( \text{CH}_3 - \text{C} - \text{O} - \text{CH}_2 - \text{CH}_3 \)
   b. \( \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{C} - \text{O} - \text{CH}_2 - \text{CH}_3 \)

17.47  a. \( \text{O} \)
   b. \( \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{C} - \text{O} - \text{CH}_2 - \text{CH}_3 \)
   c. \( \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{C} - \text{O} - \text{CH}_2 - \text{CH}_3 \)

17.49  a. methyl methanoate (methyl formate)
   b. ethyl ethanoate (ethyl acetate)
   c. ethyl propanoate (ethyl propionate)

17.51  a. \( \text{CH}_3 - \text{C} - \text{O} - \text{CH}_3 \)
   b. \( \text{H} - \text{C} - \text{O} - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{CH}_3 \)
   c. \( \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{C} - \text{O} - \text{CH}_2 - \text{CH}_3 \)
   d. \( \text{CH}_3 - \text{CH}_2 - \text{C} - \text{O} - \text{CH}_2 - \text{CH}_2 - \text{CH}_3 \)
17.53 a. methylamine  
   b. methylpropylamine  
   c. diethylmethylamine

17.55 a. CH₃—C—NH₂  
   b. CH₃—C—N—CH₂—CH₃  
   c. 

17.57 a. ethanamide (acetamide)  
   b. 2-chlorobutanamide  
   c. methanamide (formamide)

17.59 a. CH₃—CH₂—C—NH₂  
   b. CH₃—CH₂—CH₂—CH—C—NH₂  
   c. H—C—NH₂

17.61 —C—C—C—C—C—C—
                 F    F    F    F    F    F    F

17.63 a. C₉H₂₀(l) + 14O₂(g) $\xrightarrow{\Delta}$ 9CO₂(g) + 10H₂O(g) + energy  
   b. C₇H₈(l) + 9O₂(g) $\xrightarrow{\Delta}$ 7CO₂(g) + 4H₂O(g) + energy

17.65 $n\xrightarrow{\text{F}}$ —C—C—C—C—C—
                 F    F    F    F    F    F    F

17.67 a. 3,7-dimethyl-6-octenal  
   b. The en signifies that a double bond is present.  
   c. The al signifies that an aldehyde is present.

17.69 a. CH₃—CH₂—CH—CH₂—CH₂—CH₃  
   b. CH₃—CH=CH—CH₂—CH₃  
   c. CH₃—C≡C—CH₂—CH₂—CH₃

17.71 a. 2,2-dimethylbutane  
   b. 1-butyne  
   c. 2-pentene

17.73 a. ketone  
   b. alkene  
   c. ester  
   d. amine

17.75 a. ether  
   b. amine  
   c. ester  
   d. amide

17.77 a. aromatic, aldehyde  
   b. aromatic, aldehyde, alkene

17.79 a. 2,4-dichloro-1-methylbenzene  
   b. 2,4,6-trinitrophenol  
   c. 1,2-dimethyl-5-nitrobenzene

17.81 a.  
   b.  
   c.  

17.83 a. 2-butanol  
   b. 3-methyl-2-pentanol  
   c. butyl ethyl ether

17.85 a.  
   b.  
   c.  

17.87 a. 4-chlorobenzaldehyde  
   b. 3-chloropropanal  
   c. 2-chloro-3-pentanone

17.89 a.  
   b. Cl—CH₂—CH₂—C—H  
   c. CH₃—CH₂—C—CH₃

17.91 a. 3-methylbutanoic acid  
   b. methyl benzoate  
   c. ethyl propanoate

17.93 a. CH₃—CH₂—NH₂  
   b. CH₃—CH₂—CH₂—CH₂—CH₂—C—NH₂  
   c. CH₃—CH₂—N—CH₂—CH₃

17.95 a. butyldimethylamine  
   b. pentanamide  
   c. pentyamine
17.97

- 1-Pentanol
  \[ \text{CH}_3\text{CH}-\text{CH}_2\text{-CH}-\text{CH}_2\text{-OH} \]
- 2-Pentanol
  \[ \text{CH}_3\text{-CH}-\text{CH}_2\text{-CH}-\text{CH}_2\text{-OH} \]
- 3-Pentanol
  \[ \text{CH}_3\text{-CH}_2\text{-CH}-\text{CH}_2\text{-CH}_3 \]
- 2-Methyl-1-butanol
  \[ \text{HO}-\text{CH}_2\text{-CH}-\text{CH}_2\text{-CH}_3 \]
- 3-Methyl-1-butanol
  \[ \text{HO}-\text{CH}_2\text{-CH}_2\text{-CH}-\text{CH}_3 \]
- 2-Methyl-2-butanol
  \[ \text{CH}_3\text{-C}-\text{CH}_2\text{-CH}_3 \]
- 3-Methyl-2-butanol
  \[ \text{CH}_3\text{-CH}-\text{CH}-\text{CH}_3 \]
- 2,2-Dimethyl-1-propanol
  \[ \text{CH}_3\text{-C}-\text{CH}_2\text{-OH} \]

17.99

17.101

a. \[ 2\text{CH}_3\text{-CH}_2\text{-C}≡\text{CH} + 11\text{O}_2 \xrightarrow{\Delta} 8\text{CO}_2 + 6\text{H}_2\text{O} + \text{energy} \]

b. \[ \text{CH}_3\text{-CH}_2\text{-CH}≡\text{CH} + \text{CH}_3 + \text{H}_2 \xrightarrow{\text{Pt}} \text{CH}_3\text{-CH}_2\text{-CH}_2\text{-CH}_2\text{-CH}_3 \]
REBECCA LEARNED OF her family’s hypercholesterolemia after her mother died of a heart attack at 44. Her mother had a total cholesterol level of 600 mg/dL, which is three times the normal level. Rebecca learned that the lumps under her mother’s skin and around the cornea of her eyes, called xanthomas, were caused by cholesterol that was stored throughout her body. In familial hypercholesterolemia (FH), genetic mutations prevent the removal of cholesterol from the bloodstream. As a result, low-density lipoprotein (LDL) accumulates in the blood and a person with FH can develop cardiovascular disease early in life. At the lipid clinic, Rebecca’s blood tests showed that she had a total cholesterol level of 420 mg/dL and an LDL-cholesterol level of 275 mg/dL. Rebecca was diagnosed with FH. She learned that from birth, cholesterol has been accumulating throughout her arteries.

After her diagnosis, Rebecca met with Susan, a clinical lipid specialist at the lipid clinic where she informed Rebecca about managing risk factors that can lead to heart attack and stroke. Susan told Rebecca about medications she would prescribe, and then followed the results.

CAREER
Clinical Lipid Specialist
Clinical lipid specialists work with patients who have lipid disorders such as high cholesterol, high triglycerides, coronary heart disease, obesity, and FH. At a lipid clinic, a clinical lipid specialist reviews a patient’s lipid profile, which includes total cholesterol, high-density lipoprotein (HDL), and low-density lipoprotein (LDL). If a lipid disorder is identified, the lipid specialist assesses the patient’s current diet and exercise program. The lipid specialist diagnoses and determines treatment including dietary changes such as reducing salt intake, increasing the amount of fiber in the diet, and lowering the amount of fat. He or she also discusses drug therapy using lipid-lowering medications that remove LDL-cholesterol to help patients achieve and maintain good health.

Allied health professionals such as nurses, nurse practitioners, pharmacists, and dietitians are certified by a clinical lipid specialist program for specialized care of patients with lipid disorders.
18.1 Carbohydrates

**LEARNING GOAL** Classify a carbohydrate as an aldose or a ketose; draw the open-chain and Haworth structures for glucose, galactose, and fructose.

Each day, you may enjoy the polysaccharides called starches in bread and pasta. The table sugar used to sweeten cereal, tea, or coffee is sucrose, a disaccharide that consists of two simple sugars, glucose, and fructose. Carbohydrates such as table sugar, lactose in milk, and cellulose are all made of carbon, hydrogen, and oxygen. Simple sugars, which have formulas of $C_n(H_2O)_m$, were once thought to be hydrates of carbon, thus the name carbohydrate. In a series of reactions called photosynthesis, energy from the Sun is used to combine the carbon atoms from carbon dioxide ($CO_2$) and the hydrogen and oxygen atoms of water into the carbohydrate glucose.

$$6CO_2 + 6H_2O + \text{energy} \xrightarrow{\text{Photosynthesis}} C_6H_{12}O_6 + 6O_2$$

In the body, glucose is oxidized in a series of metabolic reactions known as respiration, which releases chemical energy to do work in the cells. Carbon dioxide and water are produced and returned to the atmosphere. The combination of photosynthesis and respiration is the carbon cycle, in which energy from the Sun is stored in plants by photosynthesis and made available to us when the carbohydrates in our diets are metabolized (see Figure 18.1).

**Monosaccharides**

The simplest carbohydrates are the monosaccharides. A monosaccharide cannot be split into smaller carbohydrates. A monosaccharide has a chain of three to seven carbon atoms, one in a carbonyl group and all the others are attached to hydroxyl groups (—OH). In an aldose, the carbonyl group is on the first carbon (—CHO); a ketone contains the carbonyl group on the second carbon atom as a ketone (C==O).
Monosaccharides are also classified by the number of carbon atoms. A monosaccharide with three carbon atoms is a triose, one with four carbon atoms is a tetrose; a pentose has five carbons, and a hexose contains six carbons. We can use both classification systems to indicate the aldehyde or ketone group and the number of carbon atoms. An aldopentose is a five-carbon monosaccharide that is an aldehyde; a ketohexose is a six-carbon monosaccharide that is a ketone.

**Open-Chain Structures of Some Important Monosaccharides**

Glucose, galactose, and fructose are the most important monosaccharides. They are all hexoses with the molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$. In nature and the cells of the body, their most common form is the $\text{d}$ isomer, which has the $-\text{OH}$ group attached to carbon 5 on the right side of the chain.

The most common hexose, $\text{d}-\text{glucose}$, $\text{C}_6\text{H}_{12}\text{O}_6$, an aldohexose also known as dextrose and blood sugar, is found in fruits, vegetables, corn syrup, and honey. It is a building block of the disaccharides sucrose, lactose, and maltose, and polysaccharides such as starch, cellulose, and glycogen.

$\text{d}-\text{Galactose}$ is an aldohexose that does not occur in the free form in nature. It is obtained from the disaccharide lactose, a sugar found in milk and milk products. $\text{d}-\text{Galactose}$ is important in the cellular membranes of the brain and nervous system. The only difference in the structures of $\text{d}-\text{glucose}$ and $\text{d}-\text{galactose}$ is the arrangement of the $-\text{OH}$ group on carbon 4.

In contrast to glucose and galactose, $\text{d}-\text{fructose}$ is a ketohexose. The structure of $\text{d}-\text{fructose}$ differs from $\text{d}-\text{glucose}$ at carbons 1 and 2 by the location of the carbonyl group. $\text{d}-\text{Fructose}$ is the sweetest of the carbohydrates, twice as sweet as sucrose (table sugar). This makes $\text{d}-\text{fructose}$ popular with dieters because less $\text{d}-\text{fructose}$ and, therefore, fewer calories are needed to provide a pleasant taste. $\text{d}-\text{Fructose}$ is found in fruit juices and honey.
CHEMISTRY LINK TO HEALTH

Hyperglycemia and Hypoglycemia

A doctor may order a glucose tolerance test to evaluate the body’s ability to return to normal blood glucose concentrations of 70 to 90 mg/dL of blood in response to the ingestion of a specified amount of glucose. The patient fasts for 12 h and then drinks a solution containing 75 g/dL of glucose. A blood sample is taken immediately, followed by more blood samples each half-hour for 2 h, and then every hour for a total of 5 h. If the blood glucose exceeds 200 mg/dL and remains high, hyperglycemia may be indicated. The term glyc or gluco refers to “sugar.” The prefix hyper means above or over, and hypo is below or under. Thus the blood sugar level in hyperglycemia is above normal and in hypoglycemia it is below normal.

An example of a disease that can cause hyperglycemia is type 2 diabetes, which occurs when the pancreas is unable to produce sufficient quantities of insulin. As a result, glucose levels in the body fluids can rise as high as 350 mg/dL. Symptoms of diabetes include thirst, excessive urination, and increased appetite. In older people, type 2 diabetes is sometimes a consequence of excessive weight gain.

When a person is hypoglycemic, the blood glucose level rises and then decreases rapidly to levels below 40 mg/dL. In some cases, hypoglycemia is caused by overproduction of insulin by the pancreas. Low blood glucose can cause dizziness, general weakness, and muscle tremors. A diet may be prescribed that consists of several small meals high in protein and low in carbohydrate. Some hypoglycemic patients are finding success with diets that include more complex carbohydrates rather than simple sugars.

A glucose solution is given to determine blood glucose levels.

Haworth Structures of Monosaccharides

Up until now, we have drawn the structures for monosaccharides such as d-glucose as open chains. However, the most stable form of hexoses is a six-atom ring, known as a Haworth structure. We can draw the Haworth structure for d-glucose as shown in Sample Problem 18.1.

Guide to Drawing Haworth Structures

STEP 1
Turn the open-chain structure clockwise by 90°.

STEP 2
Fold the horizontal carbon chain into a hexagon, rotate the groups on carbon 5, and bond the O on carbon 5 to carbon 1.

STEP 3
Draw the new —OH group on carbon 1 below the ring to give the α form or above the ring to give the β form.

SAMPLE PROBLEM 18.1 Drawing the Haworth Structure for d-Glucose

d-Glucose has the following open-chain structure. Draw the Haworth structure for d-glucose.

\[
\begin{align*}
\text{H} & \quad \text{C} \quad \text{O} \\
\text{H} & \quad \text{C} \quad \text{OH} \\
\text{HO} & \quad \text{C} \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{OH} \\
\text{H} & \quad \text{C} \quad \text{OH} \\
\text{CH}_2\text{OH} &
\end{align*}
\]
**TRY IT FIRST**

**SOLUTION**

**STEP 1** Turn the open-chain structure clockwise by 90°. The —H and —OH groups on the right of the vertical carbon chain are now below the horizontal carbon chain. Those on the left of the open chain are now above the horizontal carbon chain.

![Diagram of glucose structure]

**STEP 2** Fold the horizontal carbon chain into a hexagon, rotate the groups on carbon 5, and bond the O on carbon 5 to carbon 1. With carbons 2 and 3 as the base of a hexagon, move the remaining carbons upward. Rotate the groups on carbon 5 so that the —OH group is close to carbon 1. To complete the Haworth structure, draw a bond between the oxygen of the —OH group on carbon 5 to carbon 1. By convention, the carbon atoms in the ring are drawn as corners.

![Diagram of glucose structure]

**ENgage** How does a new hydroxyl group form on carbon 1 when the oxygen atom on carbon 5 reacts with carbon 1?

**STEP 3** Draw the new —OH group on carbon 1 below the ring to give the α form or above the ring to give the β form. Because the new —OH group can form above or below the plane of the Haworth structure, there are two forms of d-glucose, which differ only by the position of the —OH group at carbon 1.

![Diagram of glucose structure]
STUDY CHECK 18.1

D-Galactose is an aldohexose like D-glucose, differing only in the arrangement of the —OH group on carbon 4. Draw the Haworth structure for α-D-galactose.

In contrast to D-glucose and D-galactose, D-fructose is a ketohexose. It forms the five-atom ring when the hydroxyl group on carbon 5 reacts with the carbon of the ketone group. The new hydroxyl group is on carbon 2.

QUESTIONS AND PROBLEMS

18.1 Carbohydrates

LEARNING GOAL. Classify a carbohydrate as an aldose or a ketose; draw the open-chain and Haworth structures for glucose, galactose, and fructose.

18.1 What functional groups are found in all monosaccharides?
18.2 What is the difference between an aldose and a ketose?
18.3 What are the functional groups and number of carbons in a ketopentose?
18.4 What are the functional groups and number of carbons in an aldohexose?
18.5 What are the kind and number of atoms in the ring portion of the Haworth structure of D-glucose?
18.6 What are the kind and number of atoms in the ring portion of the Haworth structure of D-fructose?

18.7 Identify each of the following Haworth structures as the α or β form:

a. 

b. 

HOCH₂

CH₂OH

H OH

H OH

H OH

H OH

α-D-Fructose
18.8 Identify each of the following Haworth structures as the α or β form:

a.  
\[
\begin{array}{c}
\text{HO} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{CH}_2\text{OH}
\end{array}
\]

b.  
\[
\begin{array}{c}
\text{HO} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{CH}_2\text{OH}
\end{array}
\]

18.9 Classify each of the following monosaccharides as an aldopentose, aldohexose, ketopentose, or ketohexose:

a. Psicose is present in low levels in foods.

b. Lyxose is a component of bacterial lipids.

18.10 Classify each of the following monosaccharides as an aldopentose, aldohexose, ketopentose, or ketohexose:

a. A solution of xylose is given to test absorption by the intestines.

b. Tagatose, found in fruit, is similar in sweetness to sugar.

18.11 An infant with galactosemia can utilize d-glucose in milk but not d-galactose. How does the open-chain structure of d-galactose differ from that of d-glucose?

18.12 d-Fructose is the sweetest monosaccharide. How does the open-chain structure of d-fructose differ from that of d-glucose?

18.2 Disaccharides and Polysaccharides

Learning Goal Describe the monosaccharide units and linkages in disaccharides and polysaccharides.

A disaccharide is composed of two monosaccharides linked together. The most common disaccharides are maltose, lactose, and sucrose.

Maltose, or malt sugar, is a disaccharide obtained from starch and is found in germinating grains. Maltose is used in cereals, candies, and the brewing of beverages. Maltose has a glycosidic bond between two glucose molecules. To form maltose, the \(-\text{OH}\) group...
on carbon 1 in the first glucose forms a bond with the —OH group on carbon 4 in a second glucose molecule, which is designated as an α(1→4) linkage. Because the second glucose molecule in maltose has a free —OH on carbon 1, there are α and β forms of maltose.

**Lactose**, milk sugar, is a disaccharide found in milk and milk products (see **FIGURE 18.2**). It makes up 6 to 8% of human milk and about 4 to 5% of cow’s milk. Some people do not produce sufficient quantities of the enzyme needed to break down lactose, and the sugar remains undigested, causing abdominal cramps and diarrhea. In some commercial milk products, an enzyme called lactase is added to break down lactose. The bond in lactose is a β(1→4)-glycosidic bond because the β form of galactose links to the hydroxyl group on carbon 4 of glucose.

**Sucrose**, ordinary table sugar, is the most abundant disaccharide in the world. Most of the sucrose for table sugar comes from sugar cane (20% by mass) or sugar beets (15% by mass) (see **FIGURE 18.3**). Both the raw and refined forms of sugar are sucrose. Some estimates indicate that each person in the United States consumes an average of 68 kg (150 lb) of sucrose every year either by itself or in a variety of food products. Sucrose consists of an α-d-glucose and a β-d-fructose molecule joined by a bond called an α, β(1→2)-glycosidic bond.
SAMPLE PROBLEM 18.2 Glycosidic Bonds in Disaccharides

Melibiose is a disaccharide that is 30 times sweeter than sucrose.

\[ \text{Melibiose} \]

\[ \text{OH} \quad \text{OH} \quad \text{H} \quad \text{H} \]
\[ \text{H} \quad \text{H} \quad \text{CH}_2 \text{O} \]
\[ \text{OHH} \quad \text{OH} \quad \text{H} \]
\[ \text{OH} \quad \text{H} \quad \text{H} \]
\[ \text{O} \quad \text{CH}_2 \text{OH} \]

a. What are the monosaccharide units in melibiose?
b. What type of glycosidic bond links the monosaccharides?
c. Identify the structure as \( \alpha \)- or \( \beta \)-melibiose.

TRY IT FIRST

SOLUTION

a. First monosaccharide (left) When the \( \text{―OH} \) group on carbon 4 is above the plane, it is \( \text{d-galactose} \). When the \( \text{―OH} \) group on carbon 1 is below the plane, it is \( \alpha \)-d-galactose.

Second monosaccharide (right) When the \( \text{―OH} \) group on carbon 4 is below the plane, it is \( \text{d-glucose} \).

b. Type of glycosidic bond The \( \text{―OH} \) group at carbon 1 of \( \alpha \)-d-galactose bonds with the \( \text{―OH} \) group on carbon 6 of glucose, which makes it an \( \alpha(1\rightarrow6) \)-glycosidic bond.

c. Name of disaccharide The \( \text{―OH} \) group on carbon 1 of glucose is below the plane, which is \( \alpha \)-melibiose.

STUDY CHECK 18.2

Cellobiose is a disaccharide composed of two \( \text{d-glucose} \) molecules connected by a \( \beta(1\rightarrow4) \)-glycosidic bond. Draw the Haworth structure for \( \beta \)-cellobiose.

ANSWER

\[ \text{Cellobiose} \]

\[ \text{OH} \quad \text{OH} \quad \text{H} \quad \text{H} \]
\[ \text{H} \quad \text{H} \quad \text{CH}_2 \text{O} \]
\[ \text{OHH} \quad \text{OH} \quad \text{H} \]
\[ \text{OH} \quad \text{H} \quad \text{H} \]
\[ \text{O} \quad \text{CH}_2 \text{OH} \]
Although many of the monosaccharides and disaccharides taste sweet, they differ considerably in their degree of sweetness. Dietetic foods contain sweeteners that are noncarbohydrate or carbohydrates that are sweeter than sucrose. Some examples of sweeteners compared with sucrose are shown in Table 18.1.

Sucralose, which is known as Splenda, is made from sucrose by replacing some of the hydroxyl groups with chlorine atoms.

\[
\text{Sucralose (Splenda)}
\]

Aspartame, which is marketed as NutraSweet and Equal, is used in a large number of sugar-free products. It is a noncarbohydrate sweetener made of aspartate and a methyl ester of phenylalanine. It does have some caloric value, but it is so sweet that only a very small quantity is needed. However, phenylalanine, one of the breakdown products, poses a danger to anyone who cannot metabolize it properly, a condition called phenylketonuria (PKU).

Another artificial sweetener, Neotame, is a modification of the aspartame structure. The addition of a large alkyl group to the amine group prevents enzymes from breaking the amide bond between aspartate and phenylalanine. Thus, phenylalanine is not produced when Neotame is used as a sweetener. Very small amounts of Neotame are needed because it is about 10 000 times sweeter than sucrose.

Saccharin, which is marketed as Sweet’N Low, has been used as a noncarbohydrate artificial sweetener for the past 35 years.

Stevia is a sugar substitute obtained from the leaves of a plant Stevia rebaudiana. It is composed of steviol glycosides that are about 150 times sweeter than sucrose. Stevia has been used to sweeten tea and medicines in South America for 1500 years.

<table>
<thead>
<tr>
<th>Table 18.1 Relative Sweetness of Sugars and Artificial Sweeteners</th>
<th>Sweetness Relative to Sucrose (= 100)</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Monosaccharides</strong></td>
<td></td>
</tr>
<tr>
<td>Galactose</td>
<td>30</td>
</tr>
<tr>
<td>Glucose</td>
<td>75</td>
</tr>
<tr>
<td>Fructose</td>
<td>175</td>
</tr>
<tr>
<td><strong>Disaccharides</strong></td>
<td></td>
</tr>
<tr>
<td>Lactose</td>
<td>16</td>
</tr>
<tr>
<td>Maltose</td>
<td>33</td>
</tr>
<tr>
<td>Sucrose</td>
<td>100</td>
</tr>
<tr>
<td><strong>Sugar Alcohols</strong></td>
<td></td>
</tr>
<tr>
<td>Sorbitol</td>
<td>60</td>
</tr>
<tr>
<td>Maltitol</td>
<td>80</td>
</tr>
<tr>
<td>Xylitol</td>
<td>100</td>
</tr>
<tr>
<td><strong>Artificial Sweeteners (Noncarbohydrate)</strong></td>
<td></td>
</tr>
<tr>
<td>Aspartame</td>
<td>18 000</td>
</tr>
<tr>
<td>Saccharin</td>
<td>45 000</td>
</tr>
<tr>
<td>Sucralose</td>
<td>60 000</td>
</tr>
<tr>
<td>Neotame</td>
<td>1 000 000</td>
</tr>
</tbody>
</table>
**Polysaccharides**

A polysaccharide is a polymer of many monosaccharides joined together. Four biologically important polysaccharides—amylose, amylopectin, glycogen, and cellulose—are all polymers of D-glucose that differ only in the type of glycosidic bonds and the amount of branching in the molecule.

Starch, a storage form of glucose in plants, is found as insoluble granules in rice, wheat, potatoes, beans, and cereals. Starch is composed of two kinds of polysaccharides, amylose and amylopectin. Amylose, which makes up about 20% of starch, consists of α-D-glucose molecules connected by α(1→4)-glycosidic bonds in a continuous chain. A typical polymer of amylose may contain from 250 to 4000 α-D-glucose molecules. Sometimes called a straight-chain polymer, polymers of amylose are actually coiled in helical fashion.

Amylopectin, which makes up as much as 80% of plant starch, is a branched-chain polysaccharide. Like amylose, the glucose molecules are connected by α(1→4)-glycosidic bonds. However, at about every 25 glucose units, there is a branch of glucose molecules attached by an α(1→6)-glycosidic bond between carbon 1 of the branch and carbon 6 in the main chain (see FIGURE 18.4).

**FIGURE 18.4** The structure of amylose (a) is a straight-chain polysaccharide of glucose units, and amylopectin (b) is a branched chain of glucose.

What are the two types of glycosidic bonds that link glucose molecules in amylopectin?
Starches hydrolyze easily in water and acid to give smaller saccharides called \textit{dextrins}, which then hydrolyze to maltose and finally glucose. In our bodies, these complex carbohydrates are digested by the enzymes amylase (in saliva) and maltase. The glucose obtained provides about 50% of our nutritional calories.

\textbf{Glycogen}, or animal starch, is a polymer of glucose that is stored in the liver and muscle of animals. It is used in our cells at a rate that maintains the blood level of glucose and provides energy between meals. The structure of glycogen is very similar to that of amylopectin except that glycogen is more highly branched.

\textbf{Cellulose} is the major structural material of wood and plants. Cotton is almost pure cellulose. In cellulose, glucose molecules form a long unbranched chain similar to that of amylose. However, the glucose units in cellulose are linked by $\beta(1\rightarrow4)$-glycosidic bonds (see \textbf{FIGURE 18.5}).

\begin{figure}[h]
\centering
\includegraphics[width=\textwidth]{cellulose_diagram.png}
\caption{The polysaccharide cellulose is composed of glucose units linked by $\beta(1\rightarrow4)$-glycosidic bonds.}
\end{figure}

Humans have enzymes in saliva and pancreatic juices that break apart the $\alpha(1\rightarrow4)$-glycosidic bonds of starches but not the $\beta(1\rightarrow4)$-glycosidic bonds of cellulose. Thus, humans cannot digest cellulose. Animals such as horses, cows, and goats can obtain glucose from cellulose because their digestive systems contain bacteria that provide enzymes to break apart $\beta(1\rightarrow4)$-glycosidic bonds.

\textbf{SAMPLE PROBLEM 18.3 Structures of Polysaccharides}

Identify the polysaccharide described by each of the following:

\begin{itemize}
\item[a.] a polysaccharide that is stored in the liver and muscle tissues
\item[b.] an unbranched polysaccharide containing $\beta(1\rightarrow4)$-glycosidic bonds
\item[c.] a starch containing $\alpha(1\rightarrow4)$- and $\alpha(1\rightarrow6)$-glycosidic bonds
\end{itemize}
TRY IT FIRST

SOLUTION
a. glycogen  b. cellulose  c. amylepectin, glycogen

STUDY CHECK 18.3
Cellulose and amylose are both unbranched glucose polymers. How do they differ?

ANSWER
Cellulose contains glucose units connected by \( \beta(1\rightarrow 4) \)-glycosidic bonds, whereas the glucose units in amylose are connected by \( \alpha(1\rightarrow 4) \)-glycosidic bonds.

QUESTIONS AND PROBLEMS

18.2 Disaccharides and Polysaccharides

LEARNING GOAL Describe the monosaccharide units and linkages in disaccharides and polysaccharides.

18.13 For each of the following, state the monosaccharide units, the type of glycosidic bond, and the name of the disaccharide, including the \( \alpha \) or \( \beta \) form:

a. 

b.

18.14 For each of the following, state the monosaccharide units, the type of glycosidic bond, and the name of the disaccharide, including the \( \alpha \) or \( \beta \) form:

a. 

b.

Applications

18.15 Isomaltose, obtained from the breakdown of starch, has the following Haworth structure:

a. Is isomaltose a mono-, di-, or polysaccharide?
b. What are the monosaccharides in isomaltose?c. What is the glycosidic link in isomaltose?d. Is this the \( \alpha \) or \( \beta \) form of isomaltose?

18.16 Sophorose, found in certain types of beans, has the following Haworth structure:

a. Is sophorose a mono-, di-, or polysaccharide?
b. What are the monosaccharides in sophorose?c. What is the glycosidic link in sophorose?d. Is this the \( \alpha \) or \( \beta \) form of sophorose?

18.17 Identify the disaccharide that fits each of the following descriptions:

a. ordinary table sugarb. found in milk and milk products
c. also called malt sugard. contains galactose and glucose
18.3 Lipids

LEARNING GOAL Draw the condensed or line-angle structural formula for a fatty acid, a triacylglycerol, and the products of hydrogenation or saponification. Identify the steroid nucleus.

Lipids are a family of biomolecules that have the common property of being soluble in organic solvents but not very soluble in water. The word lipid comes from the Greek word lipos, meaning “fat” or “lard.” Within the lipid family, there are certain structures that distinguish the different types of lipids. Lipids that contain fatty acids include waxes and triacylglycerols, commonly known as fats and oils, which are esters of glycerol and fatty acids. Lipids that are steroids do not contain fatty acids but are characterized by the steroid nucleus of four fused carbon rings.

Lipids are naturally occurring compounds in cells and tissues, which are soluble in organic solvents and not in water.

Fatty Acids

A fatty acid contains a long unbranched chain of carbon atoms, usually 12 to 20, with a carboxylic acid group at one end. The long carbon chain makes fatty acids insoluble in water. An example is lauric acid, a 12-carbon acid found in coconut oil, which has a structure that can be drawn in several forms as follows:

\[
\text{CH}_3\text{--(CH}_2\text{)}_{10}\text{--C--OH} \quad \text{CH}_3\text{--(CH}_2\text{)}_{10}\text{--COOH}
\]

\[
\text{OH}
\]

Condensed structural formulas

Line-angle structural formula
Saturated fatty acids, such as lauric acid, contain only single bonds between carbons. Monounsaturated fatty acids have one double bond in the carbon chain, and polyunsaturated fatty acids have two or more double bonds. **TABLE 18.2** lists some of the typical fatty acids in lipids.

**TABLE 18.2** Structures and Melting Points of Common Fatty Acids

<table>
<thead>
<tr>
<th>Name</th>
<th>Carbon Atoms: Double Bonds</th>
<th>Source</th>
<th>Melting Point (°C)</th>
<th>Structural Formulas</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Saturated Fatty Acids</strong></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Lauric acid</td>
<td>12:0</td>
<td>Coconut</td>
<td>44</td>
<td>$\text{CH}_3 \text{-(CH}<em>2\text{)}</em>{10}\text{-COOH}$</td>
</tr>
<tr>
<td>Myristic acid</td>
<td>14:0</td>
<td>Nutmeg</td>
<td>55</td>
<td>$\text{CH}_3 \text{-(CH}<em>2\text{)}</em>{12}\text{-COOH}$</td>
</tr>
<tr>
<td>Palmitic acid</td>
<td>16:0</td>
<td>Palm</td>
<td>63</td>
<td>$\text{CH}_3 \text{-(CH}<em>2\text{)}</em>{14}\text{-COOH}$</td>
</tr>
<tr>
<td>Stearic acid</td>
<td>18:0</td>
<td>Animal fat</td>
<td>69</td>
<td>$\text{CH}_3 \text{-(CH}<em>2\text{)}</em>{16}\text{-COOH}$</td>
</tr>
<tr>
<td><strong>Monounsaturated Fatty Acids</strong></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Palmitoleic acid</td>
<td>16:1</td>
<td>Butter</td>
<td>0</td>
<td>$\text{CH}_3 \text{-(CH}<em>2\text{)}</em>{3}\text{CH=CH-(CH}<em>2\text{)}</em>{7}\text{-COOH}$</td>
</tr>
<tr>
<td>Oleic acid</td>
<td>18:1</td>
<td>Olive, pecan, grapeseed</td>
<td>14</td>
<td>$\text{CH}_3 \text{-(CH}<em>2\text{)}</em>{7}\text{CH=CH-(CH}<em>2\text{)}</em>{7}\text{-COOH}$</td>
</tr>
<tr>
<td><strong>Polyunsaturated Fatty Acids</strong></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Linoleic acid</td>
<td>18:2</td>
<td>Soybean, sunflower</td>
<td>$-5$</td>
<td>$\text{CH}_3 \text{-(CH}<em>2\text{)}</em>{4}\text{CH=CH-CH}_2\text{CH=CH-(CH}<em>2\text{)}</em>{7}\text{-COOH}$</td>
</tr>
<tr>
<td>Linolenic acid</td>
<td>18:3</td>
<td>Corn</td>
<td>$-11$</td>
<td>$\text{CH}_3 \text{-CH}_2\text{CH=CH-CH}_2\text{CH=CH-CH}_2\text{CH=CH-(CH}<em>2\text{)}</em>{7}\text{-COOH}$</td>
</tr>
</tbody>
</table>

**Cis–Trans Isomers**

Compounds with double bonds can be drawn as two structures known as *cis* and *trans* isomers. In 2-butene, which has a double bond, the *cis* isomer has the carbon groups on the same side of the double bond, but the *trans* isomer has the carbon groups on opposite sides of the double bond.
We can draw the unsaturated fatty acid oleic acid as cis and trans isomers using its line-angle structural formula. The cis structure is the more prevalent isomer found in naturally occurring unsaturated fatty acids. In the cis isomer, the carbon chain has a “kink” at the double bond site. In contrast, the trans isomer, called elaidic acid, is a straight chain without a kink at the double bond site. Saturated fatty acids fit close together in a regular pattern, which allows strong attractions to occur between the carbon chains. However, unsaturated fatty acids have irregular shape, which leads to fewer interactions between carbon chains. Thus, saturated fatty acids have high melting points because more energy is required to separate the fatty acid molecules.

Almost all naturally occurring unsaturated fatty acids have one or more cis double bonds.

**CHEMISTRY LINK TO HEALTH**

**Trans Fatty Acids and Hydrogenation**

During the early 1900s, margarine became a popular replacement for highly saturated fats such as butter and lard. Margarine is produced by partially hydrogenating the unsaturated fats in vegetable oils such as safflower oil, corn oil, canola oil, cottonseed oil, and sunflower oil.

**Hydrogenation and Trans Fats**

In vegetable oils, the unsaturated fats usually contain cis double bonds. As hydrogenation occurs, double bonds are converted to single bonds. However, a small number of the cis double bonds are converted to trans double bonds. If the label on a product states that the oils have been “partially hydrogenated,” that product will contain trans fatty acids.

The concern about trans fatty acids is that their altered structure may make them behave like saturated fatty acids in the body. Several studies reported that trans fatty acids raise the levels of LDL-cholesterol, low-density lipoproteins containing cholesterol that can accumulate in the arteries. Some studies also report that trans fatty acids lower HDL-cholesterol, high-density lipoproteins that carry cholesterol to the liver to be excreted.

Foods containing trans fatty acids include baked goods, stick margarine, soft margarine, cookies, crackers, and vegetable shortening. The American Heart Association recommends that margarine should have...
The best advice may be to reduce total fat in the diet by using fats and oils sparingly, cooking with little or no fat, substituting olive oil or canola oil for other oils, and limiting the use of coconut oil and palm oil, which are high in saturated fatty acids.
Waxes

Waxes are found in many plants and animals. A wax is an ester of a long-chain fatty acid and a long-chain alcohol. The formulas of some common waxes are given in Table 18.3. Beeswax obtained from honeycombs and carnauba wax from palm trees are used to give a protective coating to furniture, cars, and floors. Jojoba wax is used in making candles and cosmetics such as lipstick. Lanolin, a mixture of waxes obtained from wool, is used in hand and facial lotions to aid retention of water, which softens the skin.

<table>
<thead>
<tr>
<th>Type</th>
<th>Condensed Structural Formula</th>
<th>Source</th>
<th>Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Beeswax</td>
<td>( \text{CH}_3\left(\text{CH}<em>2\right)</em>{14}\text{C-O-(CH}<em>2\right)</em>{20}\text{CH}_3 )</td>
<td>Honeycomb</td>
<td>Candles, shoe polish, wax paper</td>
</tr>
<tr>
<td>Carnauba wax</td>
<td>( \text{CH}_3\left(\text{CH}<em>2\right)</em>{24}\text{C-O-(CH}<em>2\right)</em>{20}\text{CH}_3 )</td>
<td>Brazilian palm tree</td>
<td>Waxes for furniture, cars, floors, shoes</td>
</tr>
<tr>
<td>Jojoba wax</td>
<td>( \text{CH}_3\left(\text{CH}<em>2\right)</em>{18}\text{C-O-(CH}<em>2\right)</em>{16}\text{CH}_3 )</td>
<td>Jojoba bush</td>
<td>Candles, soaps, cosmetics</td>
</tr>
</tbody>
</table>

Fats and Oils: Triacylglycerols

In the body, fatty acids are stored as fats and oils known as triacylglycerols. These substances, also called triglycerides, are triesters of glycerol (a trihydroxy alcohol) and fatty acids. In a triacylglycerol, three hydroxyl groups on glycerol form ester bonds with the carboxyl groups of three fatty acids. For example, glycerol and three molecules of stearic acid form glycercyl tristearate (tristearin).
Triacylglycerols are the major form of energy storage for animals. Animals that hibernate eat large quantities of plants, seeds, and nuts that are high in calories. Prior to hibernation, these animals, such as polar bears, gain as much as 14 kg a week. As the external temperature drops, the animal goes into hibernation. The body temperature drops to nearly freezing, and there is a dramatic reduction in cellular activity, respiration, and heart rate. Animals that live in extremely cold climates will hibernate for 4 to 7 months. During this time, stored fat is the only source of energy.

**Melting Points of Fats and Oils**

A **fat** is a triacylglycerol that is solid at room temperature, such as fats in meat, whole milk, butter, and cheese. An **oil** is a triacylglycerol that is usually liquid at room temperature and is usually obtained from a plant source (see FIGURE 18.6).

The amounts of saturated, monounsaturated, and polyunsaturated fatty acids in some typical fats and oils are shown in **FIGURE 18.7**. Saturated fatty acids have higher melting points than unsaturated fatty acids because they pack together more tightly. Animal fats usually contain more saturated fatty acids than do vegetable oils. Therefore the melting points of animal fats are higher than those of vegetable oils.
Triacylglycerols are used to thicken creams and lotions.

**SAMPLE PROBLEM 18.5 Drawing the Structure for a Triacylglycerol**

Draw the condensed structural formula for glyceryl tripalmitoleate (tripalmitolein), which is used in cosmetic creams and lotions.

**TRY IT FIRST**

**SOLUTION**

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>glyceryl tripalmitoleate (tripalmitolein)</td>
<td>condensed structural formula</td>
<td>ester of glycerol and three palmitoleic acid molecules</td>
</tr>
</tbody>
</table>

**STEP 1** Draw the condensed structural formulas for glycerol and the fatty acids.

![Glycerol and Palmitoleic acid structures]

**STEP 2** Form ester bonds between the hydroxyl groups on glycerol and the carboxyl groups on each fatty acid.
**STEP 2** Form ester bonds between the hydroxyl groups on glycerol and the carboxyl groups on each fatty acid.

\[
\begin{align*}
\text{CH}_2 - \overset{\text{O}}{\text{-C}} - (\text{CH}_2)_7 - \overset{\text{O}}{\text{CH}} = \overset{\text{CH}}{\text{CH}} - (\text{CH}_2)_5 - \overset{\text{O}}{\text{CH}_3} \\
\text{CH} - \overset{\text{O}}{\text{-C}} - (\text{CH}_2)_7 - \overset{\text{O}}{\text{CH}} = \overset{\text{CH}}{\text{CH}} - (\text{CH}_2)_5 - \overset{\text{O}}{\text{CH}_3} \\
\text{CH}_2 - \overset{\text{O}}{\text{-C}} - (\text{CH}_2)_7 - \overset{\text{O}}{\text{CH}} = \overset{\text{CH}}{\text{CH}} - (\text{CH}_2)_5 - \overset{\text{O}}{\text{CH}_3}
\end{align*}
\]

**Glyceryl tripalmitoleate (tripalmitolein)**

**STUDY CHECK 18.5**

Draw the line-angle structural formula for the triacylglycerol containing three molecules of myristic acid (14:0).

**ANSWER**

\[
\begin{align*}
\text{CH}_2 - \overset{\text{O}}{\text{-C}} - (\text{CH}_2)_7 - \overset{\text{O}}{\text{CH}} = \overset{\text{CH}}{\text{CH}} - (\text{CH}_2)_5 - \overset{\text{O}}{\text{CH}_3} \\
\text{CH} - \overset{\text{O}}{\text{-C}} - (\text{CH}_2)_7 - \overset{\text{O}}{\text{CH}} = \overset{\text{CH}}{\text{CH}} - (\text{CH}_2)_5 - \overset{\text{O}}{\text{CH}_3} \\
\text{CH}_2 - \overset{\text{O}}{\text{-C}} - (\text{CH}_2)_7 - \overset{\text{O}}{\text{CH}} = \overset{\text{CH}}{\text{CH}} - (\text{CH}_2)_5 - \overset{\text{O}}{\text{CH}_3}
\end{align*}
\]

**Reactions of Triacylglycerols**

The double bonds in unsaturated fatty acids react with hydrogen to produce saturated fatty acids. For example, when hydrogen is added to glyceryl trioleate (triolein) using a nickel catalyst, the product is the saturated fat glyceryl tristearate (tristearin).

**Core Chemistry Skill**

**Drawing the Products for the Hydrogenation and Saponification of a Triacylglycerol**

**Saponification** occurs when a fat is heated with a strong base such as sodium hydroxide to give glycerol and the sodium salts of the fatty acids, which is soap. When NaOH is used, a solid soap is produced that can be molded into a desired shape; KOH produces a

\[
\begin{align*}
\text{Fat or oil} & \quad + \quad \text{strong base} \quad \overset{\text{Heat}}{\longrightarrow} \quad \text{glycerol} \quad + \quad \text{salts of fatty acids (soap)} \\
\text{CH}_2 - \overset{\text{O}}{\text{-C}} - (\text{CH}_2)_7 - \overset{\text{O}}{\text{CH}} = \overset{\text{CH}}{\text{CH}} - (\text{CH}_2)_5 - \overset{\text{O}}{\text{CH}_3} & \quad + \quad 3\text{NaOH} \quad \overset{\text{Heat}}{\longrightarrow} \quad \text{CH}_2 - \overset{\text{OH}}{\text{-C}} - (\text{CH}_2)_7 - \overset{\text{O}}{\text{CH}} = \overset{\text{CH}}{\text{CH}} - (\text{CH}_2)_5 - \overset{\text{O}}{\text{CH}_3} \\
\text{Glyceryl tripalmitate (tripalmitin)} & \quad + \quad \text{Glycerol} \quad + \quad 3\text{Sodium palmitate (soap)}
\end{align*}
\]
softer, liquid soap. Oils that are polyunsaturated produce softer soaps. Names like “coconut soap” or “avocado shampoo” tell you the sources of the oil used in the reaction.

**Steroids: Cholesterol**

**Steroids** are compounds containing the steroid nucleus, which consists of four carbon rings fused together. Although they are lipids, steroids do not contain fatty acids.

Attaching other atoms and groups of atoms to the steroid nucleus forms a wide variety of steroid compounds. **Cholesterol**, which is one of the most important and abundant steroids in the body, is a sterol because it contains an oxygen atom as a hydroxyl group (—OH). Like many steroids, cholesterol has methyl groups, a double bond, and a carbon side chain. In other steroids, the hydroxyl group is replaced by a carbonyl group (C=O). Cholesterol is obtained from eating meats, milk, and eggs, and it is also synthesized by the liver. There is no cholesterol in vegetable and plant products.

**Cholesterol in the Body**

If a diet is high in cholesterol, the liver produces less cholesterol. A typical daily American diet includes 400 to 500 mg of cholesterol, one of the highest in the world. The American Heart Association has recommended that we consume no more than 300 mg of cholesterol a day. Researchers suggest that saturated fats and cholesterol are associated with diseases such as diabetes; cancers of the breast, pancreas, and colon; and atherosclerosis. In atherosclerosis, deposits of a protein–lipid complex (plaque) accumulate in the coronary blood vessels, restricting the flow of blood to the tissue and causing necrosis of the tissue (see Figure 18.8). In the heart, plaque accumulation could result in a myocardial infarction (heart attack). Other factors that may also increase the risk of heart disease are family history, lack of exercise, smoking, obesity, diabetes, gender, and age.

**QUESTIONS AND PROBLEMS**

**18.3 Lipids**

**LEARNING GOAL** Draw the condensed or line-angle structural formula for a fatty acid, a triacylglycerol, and the products of hydrogenation or saponification. Identify the steroid nucleus.

18.21 Classify each of the following fatty acids as saturated, monounsaturated, or polyunsaturated (see Table 18.2):
   a. lauric acid  b. linolenic acid  c. stearic acid

18.22 Classify each of the following fatty acids as saturated, monounsaturated, or polyunsaturated (see Table 18.2):
   a. linoleic acid  b. palmitoleic acid  c. myristic acid

18.23 Caprylic acid is an 8-carbon saturated fatty acid that occurs in coconut oil (10%) and palm kernel oil (4%). Draw the condensed structural formula for glyceryl tricaprylate (tricaprylin).

18.24 Draw the condensed structural formula for glyceryl trilaurate (trilaurin).

**Applications**

18.25 Safflower oil is called a polyunsaturated oil, whereas olive oil is a monounsaturated oil. Explain.

18.26 Why does olive oil have a lower melting point than butter fat?

18.27 Write the balanced chemical equation for the hydrogenation of glyceryl tripalmitoleate, a fat containing glycerol and three palmitoleic acid molecules.

18.28 Write the balanced chemical equation for the hydrogenation of glyceryl trilinolenate, a fat containing glycerol and three linolenic acid molecules.

18.29 Write the balanced chemical equation for the NaOH saponification of glyceryl trimyristate, a fat containing glycerol and three myristic acid molecules.

18.30 Write the balanced chemical equation for the NaOH saponification of glyceryl trioelate, a fat containing glycerol and three oleic acid molecules.

18.31 Draw the structure for the steroid nucleus.

18.32 What are the functional groups on the cholesterol molecule?
18.4 Amino Acids and Proteins

LEARNING GOAL Describe protein functions and draw structures for amino acids and peptides.

Proteins perform different functions in the body. Some proteins form structural components such as cartilage, muscles, hair, and nails. Wool, silk, feathers, and horns in animals are made of proteins. Proteins that function as enzymes regulate biological reactions such as digestion and cellular metabolism. Other proteins, such as hemoglobin and myoglobin, transport oxygen in the blood and muscle. TABLE 18.4 gives examples of proteins that are classified by their functions in biological systems.

<table>
<thead>
<tr>
<th>Class of Protein</th>
<th>Function</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>Structural</td>
<td>Provide structural components</td>
<td>Collagen is in tendons and cartilage. Keratin is in hair, skin, wool, horns, and nails.</td>
</tr>
<tr>
<td>Contractile</td>
<td>Make muscles move</td>
<td>Myosin and actin contract muscle fibers.</td>
</tr>
<tr>
<td>Transport</td>
<td>Carry essential substances throughout the body</td>
<td>Hemoglobin transports oxygen. Lipoproteins transport lipids.</td>
</tr>
<tr>
<td>Storage</td>
<td>Store nutrients</td>
<td>Casein stores protein in milk. Ferritin stores iron in the spleen and liver.</td>
</tr>
<tr>
<td>Hormone</td>
<td>Regulate body metabolism and the nervous system</td>
<td>Insulin regulates blood glucose level. Growth hormone regulates body growth.</td>
</tr>
<tr>
<td>Enzyme</td>
<td>Catalyze biochemical reactions in the cells</td>
<td>Sucrase catalyzes the hydrolysis of sucrose. Trypsin catalyzes the hydrolysis of proteins.</td>
</tr>
<tr>
<td>Protection</td>
<td>Recognize and destroy foreign substances</td>
<td>Immunoglobulins stimulate immune responses.</td>
</tr>
</tbody>
</table>

Amino Acids

Proteins are composed of molecular building blocks called amino acids. Every amino acid, which is ionized in biological environments, has a central alpha carbon atom (α carbon) bonded to an ammonium group (—NH₃⁺), a carboxylate group (—COO⁻), a hydrogen atom, and an R group. The differences in the 20 α-amino acids present in human proteins are due to the unique characteristics of the R groups.

Ionized Structure of Alanine, an α-Amino Acid

We classify amino acids according to their R groups, which determine their properties in aqueous solution. The nonpolar amino acids have hydrogen, alkyl, or aromatic R groups, which make them hydrophobic (“water fearing”). A polar amino acid interacts with water because it has a polar R group, which is hydrophilic (“water loving”). Neutral polar amino acids contain hydroxyl (—OH), thiol (—SH), or amide (—CONH₂) R groups. The R group of a polar acidic amino acid contains a carboxylate group (—COO⁻). The R group of a polar basic amino acid contains an ammonium group (—NH₃⁺). The ionized structures and common names of the 20 α-amino acids commonly found in proteins with their R groups highlighted in tan, and their three-letter and one-letter abbreviations, are listed in TABLE 18.5.
### TABLE 18.5 Structures, Names, and Abbreviations of 20 Common Amino Acids at Physiological pH (7.4)

#### Nonpolar Amino Acids (hydrophobic)

<table>
<thead>
<tr>
<th>Structure</th>
<th>Name</th>
<th>Abbreviation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Glycine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Gly (G)</td>
</tr>
<tr>
<td>Alanine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Ala (A)</td>
</tr>
<tr>
<td>Valine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Val (V)</td>
</tr>
<tr>
<td>Leucine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Leu (L)</td>
</tr>
<tr>
<td>Isoleucine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Ile (I)</td>
</tr>
<tr>
<td>Phenylalanine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Phe (F)</td>
</tr>
<tr>
<td>Methionine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Met (M)</td>
</tr>
<tr>
<td>Proline</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Pro (P)</td>
</tr>
<tr>
<td>Tryptophan</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Trp (W)</td>
</tr>
</tbody>
</table>

#### Polar Amino Acids (hydrophilic)

##### Amino Acids with Neutral R Groups

<table>
<thead>
<tr>
<th>Structure</th>
<th>Name</th>
<th>Abbreviation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Serine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Ser (S)</td>
</tr>
<tr>
<td>Threonine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Thr (T)</td>
</tr>
<tr>
<td>Tyrosine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Tyr (Y)</td>
</tr>
<tr>
<td>Cysteine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Cys (C)</td>
</tr>
<tr>
<td>Asparagine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Asn (N)</td>
</tr>
<tr>
<td>Glutamine</td>
<td>H(_3)N–C–COO(^-)</td>
<td>Gln (Q)</td>
</tr>
</tbody>
</table>

##### Amino Acids with Charged R Groups

**Acidic (negative charge)**

- Aspartate (Asp, D)
- Glutamate (Glu, E)

**Basic (positive charge)**

- Histidine (His, H)
- Lysine (Lys, K)
- Arginine (Arg, R)
A summary of the classification of amino acids follows:

<table>
<thead>
<tr>
<th>Type of Amino Acid</th>
<th>Number</th>
<th>Type of R Groups</th>
<th>Interaction with Water</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nonpolar</td>
<td>9</td>
<td>Nonpolar</td>
<td>Hydrophobic</td>
</tr>
<tr>
<td>Polar, neutral</td>
<td>6</td>
<td>Contain O and S atoms, but no charge</td>
<td>Hydrophilic</td>
</tr>
<tr>
<td>Polar, acidic</td>
<td>2</td>
<td>Contain carboxylate groups, negative charge</td>
<td>Hydrophilic</td>
</tr>
<tr>
<td>Polar, basic</td>
<td>3</td>
<td>Contain ammonium groups, positive charge</td>
<td>Hydrophilic</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 18.6 Structural Formulas of Amino Acids**

Draw the condensed structural formula and give the abbreviations for serine. Serine (R = —CH₂—OH)

**TRY IT FIRST**

**SOLUTION**

The structure of a specific amino acid is drawn by attaching the side group (R) to the central carbon atom of the general structure of an amino acid.

![Condensed structural formula of serine](image)

The abbreviations for serine are Ser and S.

**STUDY CHECK 18.6**

Classify serine in Sample Problem 18.6 as polar or nonpolar and hydrophobic or hydrophilic.

**ANSWER**

polar, hydrophilic

**CHEMISTRY LINK TO HEALTH**

**Essential Amino Acids**

Of the 20 amino acids used to build the proteins in the body, 11 can be synthesized in the body. The other nine amino acids, listed in **TABLE 18.6**, are essential amino acids that cannot be synthesized and must be obtained from the proteins in the diet. In addition to the essential amino acids required by adults, infants and growing children also require arginine, cysteine, and tyrosine.

Complete proteins, which contain all of the essential amino acids, are found in animal products such as eggs, milk, meat, fish, and poultry. However, gelatin and plant proteins such as grains, beans, and nuts are incomplete proteins because they are deficient in one or more of the essential amino acids. Diets that rely on plant foods for protein must contain a variety of protein sources to obtain all the essential amino acids. Some examples of complementary protein sources include rice and beans, or a peanut butter sandwich on whole grain bread. Rice and wheat contain the amino acid methionine that is deficient in beans and peanuts, whereas beans and peanuts are rich in lysine that is lacking in grains (see **TABLE 18.7**).
Complete proteins such as eggs, milk, meat, and fish contain all of the essential amino acids. Incomplete proteins from plants such as grains, beans, and nuts are deficient in one or more essential amino acids.

**Peptides**

A *peptide bond* is an amide bond that forms when the $\text{--COO}^-$ group of one amino acid reacts with the $\text{--NH}_3^+$ group of the next amino acid. We can write the formation of the dipeptide glycylalanine (Gly–Ala, GA) between glycine and alanine. In this dipeptide, glycine, which is written on the left with a free $\text{--NH}_3^+$ group, is the *N*-terminal amino acid or *N*-terminus. Alanine, which is written on the right with a free $\text{--COO}^-$ group, is called the *C*-terminal amino acid or *C*-terminus. In the name of a peptide, each amino acid beginning from the N-terminus has the *ine, an, or ate* replaced by *yl*. The last amino acid at the C-terminus uses its full name.

<table>
<thead>
<tr>
<th>Food Source</th>
<th>Amino Acid Deficiency</th>
</tr>
</thead>
<tbody>
<tr>
<td>Wheat, rice, oats</td>
<td>Lysine</td>
</tr>
<tr>
<td>Corn</td>
<td>Lysine, tryptophan</td>
</tr>
<tr>
<td>Beans</td>
<td>Methionine, tryptophan</td>
</tr>
<tr>
<td>Peas, peanuts</td>
<td>Methionine</td>
</tr>
<tr>
<td>Almonds, walnuts</td>
<td>Lysine, tryptophan</td>
</tr>
<tr>
<td>Soy</td>
<td>Methionine</td>
</tr>
</tbody>
</table>

**SAMPLE PROBLEM 18.7 Drawing a Peptide**

Draw the structure and give the name for the tripeptide Gly–Ser–Met.

**TRY IT FIRST**

**SOLUTION**

**Guide to Drawing a Peptide**

**STEP 1**

Draw the structures for each amino acid in the peptide, starting with the N-terminus.

**STEP 2**

Remove the O atom from the carboxylate group and two H atoms from the ammonium group in the adjacent amino acid. Use peptide bonds to connect the amino acids.
**STEP 2** Remove the O atom from the carboxylate group and two H atoms from the ammonium group in the adjacent amino acid. Use peptide bonds to connect the amino acids. Repeat this process until the C-terminus is reached.

![Amino acid structure](image)

The tripeptide is named by replacing the last syllable of each amino acid name with *yl*, starting with the N-terminus. The C-terminus retains its complete amino acid name.

**N-terminus**
- glycine is named glycyl
- serine is named seryl

**C-terminus**
- methionine keeps its full name

The tripeptide is named glycylserylmethionine.

**STUDY CHECK 18.7**

Draw the structure and give the name for Phe–Thr, a section in glucagon, which is a peptide hormone that increases blood glucose levels.

**ANSWER**

![Peptide structure](image)

**QUESTIONS AND PROBLEMS**

**LEARNING GOAL** Describe protein functions and draw structures for amino acids and peptides.

18.33 Draw the ionized form for each of the following amino acids:
   - a. glycine
   - b. threonine
   - c. phenylalanine

18.34 Draw the ionized form for each of the following amino acids:
   - a. tyrosine
   - b. leucine
   - c. methionine

18.35 Classify each of the amino acids in problem 18.33 as nonpolar or polar. If polar, indicate if the R group is neutral, acidic, or basic. Indicate if each would be hydrophobic or hydrophilic.

18.36 Classify each of the amino acids in problem 18.34 as nonpolar or polar. If polar, indicate if the R group is neutral, acidic, or basic. Indicate if each would be hydrophobic or hydrophilic.

18.37 Give the name of the amino acid represented by each of the following abbreviations:
   - a. Ala
   - b. V
   - c. Lys
   - d. C

18.38 Give the name of the amino acid represented by each of the following abbreviations:
   - a. Trp
   - b. M
   - c. Pro
   - d. G
18.5 Protein Structure

**LEARNING GOAL** Identify the levels of structure of a protein.

A **protein** is a polypeptide of 50 or more amino acids that has biological activity. Each protein in our cells has a unique sequence of amino acids that determines its threedimensional structure and biological function.

**Primary Structure**

The **primary structure** of a protein is the particular sequence of amino acids held together by peptide bonds. The first protein to have its primary structure determined was insulin, which was accomplished by Frederick Sanger in 1953. Since that time, scientists have determined the amino acid sequences of thousands of proteins. Insulin is a hormone that regulates the glucose level in the blood. In the primary structure of human insulin, there are two polypeptide chains. In chain A, there are 21 amino acids, and in chain B there are 30 amino acids. The polypeptide chains are held together by **disulfide bonds** formed by the thiol groups of the cysteine amino acids in each of the chains (see **Figure 18.9**). Today, human insulin, for the treatment of diabetes, with this exact structure is produced in large quantities through genetic engineering.

**Secondary Structure**

The **secondary structure** of a protein describes the structure that forms when amino acids form hydrogen bonds between the atoms in the backbone within a single polypeptide chain or between polypeptide chains. The most common types of secondary structures are the **alpha helix** and the **beta-pleated sheet**.

In an alpha helix (α helix), hydrogen bonds form between each N–H group and the oxygen of a C=O group in the next turn of the α helix (see **Figure 18.10**). Because there are many hydrogen bonds along the peptide, it has a helical or coiled shape.

In the secondary structure known as the beta-pleated sheet (β-pleated sheet), hydrogen bonds hold polypeptide chains together side by side. The hydrogen bonds holding the sheets tightly in place account for the strength and durability of proteins such as silk (see **Figure 18.11**).

**Collagen**

**Collagen**, the most abundant protein in the body, makes up as much as one-third of all the protein in vertebrates. It is found in connective tissue, blood vessels, skin, tendons, ligaments, the cornea of the eye, and cartilage. The strong structure of collagen is a result of three polypeptides woven together by hydrogen bonds to form a **triple helix**. When several triple helices wrap together, they form the fibrils that make up connective tissues and tendons. In a young person, collagen is elastic. As a person ages, additional bonds form between the fibrils, which make collagen less elastic. Cartilage and tendons become more brittle, and wrinkles are seen in the skin.
The α (alpha) helix acquires a coiled shape from hydrogen bonds between the N—H of the peptide bond in one turn of the polypeptide and the C=O of the peptide bond in the next turn.

What are partial charges of the H in N—H and the O in C=O that allow for hydrogen bonding?

Collagen fibers are triple helices of polypeptide chains held together by hydrogen bonds.

In a β (beta)-pleated sheet secondary structure, hydrogen bonds form between side-by-side sections of the peptide chains.

How do the hydrogen bonds differ between a β-pleated sheet and an α helix?
Tertiary Structure

The tertiary structure of a protein involves attractions and repulsions between the R groups of the amino acids in the polypeptide chain. It is the unique three-dimensional shape of the tertiary structure that determines the biological function of the molecule. Table 18.8 lists the interactions that stabilize the tertiary structures of proteins.

<table>
<thead>
<tr>
<th>Interaction</th>
<th>Nature of Bonding</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrophobic</td>
<td>Interactions between nonpolar groups</td>
</tr>
<tr>
<td>Hydrophilic</td>
<td>Attractions between polar groups and water</td>
</tr>
<tr>
<td>Salt bridges (ionic bonds)</td>
<td>Ionic interactions between acidic and basic amino acids</td>
</tr>
<tr>
<td>Hydrogen bonds</td>
<td>Attractions between H and O or N</td>
</tr>
<tr>
<td>Disulfide bonds</td>
<td>Strong covalent links between sulfur atoms of two cysteine amino acids</td>
</tr>
</tbody>
</table>

Why is the interaction between arginine and aspartate in a tertiary structure called a salt bridge?

Interactions between amino acid R groups fold a polypeptide into a specific three-dimensional shape called its tertiary structure.

**SAMPLE PROBLEM 18.8 Interactions That Stabilize Tertiary Structures**

What interaction would you expect between the following amino acids?

<table>
<thead>
<tr>
<th>ANALYZE THE PROBLEM</th>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>pairs of amino acids</td>
<td>interaction between amino acids</td>
<td>specific R group</td>
</tr>
</tbody>
</table>

a. cysteine and cysteine
b. glutamate and lysine

**TRY IT FIRST**

**SOLUTION**

a. Two cysteines, each containing $-SH$ in their R group, will form a disulfide bond.

b. An ionic bond (salt bridge) can form by the interaction of the $-COO^-$ in the R group of glutamate and the $-NH_3^+$ in the R group of lysine.
**STUDY CHECK 18.8**

What interaction would you expect between valine and leucine in the tertiary structure of a protein?

**ANSWER**

Both valine and leucine have nonpolar R groups and would have a hydrophobic interaction.

Myoglobin, a single polypeptide chain of 153 amino acids, stores oxygen in skeletal muscle. It forms a compact tertiary structure that contains a pocket of amino acids and a heme group that binds and stores one oxygen \((O_2)\) molecule (see **FIGURE 18.12**).

**Quaternary Structure**

When a biologically active protein consists of two or more polypeptide subunits, the structural level is referred to as a *quaternary structure*. Hemoglobin, a protein that transports oxygen in blood, consists of four polypeptide chains or subunits. The subunits are held together in the quaternary structure by the same kinds of interactions that stabilize the tertiary structures. In the hemoglobin molecule, the quaternary structure of hemoglobin can bind and transport four molecules of oxygen. **TABLE 18.9** and **FIGURE 18.13** summarize the structural levels of proteins.

![Summary of Structural Levels in Proteins](image)

**FIGURE 18.12**  ▶ The ribbon model represents the tertiary structure of the polypeptide chain of myoglobin that binds oxygen.

**TABLE 18.9**  Summary of Structural Levels in Proteins

<table>
<thead>
<tr>
<th>Structural Level</th>
<th>Characteristics</th>
</tr>
</thead>
<tbody>
<tr>
<td>Primary</td>
<td>Peptide bonds join amino acids in a specific sequence in a polypeptide.</td>
</tr>
<tr>
<td>Secondary</td>
<td>Hydrogen bonds along or between peptide chains form α helix, or a β-pleated sheet.</td>
</tr>
<tr>
<td>Tertiary</td>
<td>A protein folds into a compact, three-dimensional shape stabilized by interactions between R groups of amino acids.</td>
</tr>
<tr>
<td>Quaternary</td>
<td>Two or more protein subunits combine to form a biologically active protein stabilized by the same interactions as in the tertiary structure.</td>
</tr>
</tbody>
</table>

**FIGURE 18.13**  ▶ Proteins consist of (a) primary, (b) secondary, (c) tertiary, and often (d) quaternary structural levels.

**Q** What is the difference between a tertiary structure and a quaternary structure?
SAMPLE PROBLEM 18.9 Identifying Protein Structure

Indicate whether the following interactions are responsible for the primary, secondary, tertiary, or quaternary structures of proteins:

a. disulfide bonds between portions of a protein chain
b. peptide bonds that form a chain of amino acids
c. hydrophobic interactions

TRY IT FIRST

SOLUTION

Analyze the Problem

<table>
<thead>
<tr>
<th>Given</th>
<th>Need</th>
<th>Connect</th>
</tr>
</thead>
<tbody>
<tr>
<td>interactions between</td>
<td>structural level</td>
<td>identify characteristics</td>
</tr>
<tr>
<td>amino acids</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

a. Disulfide bonds help to stabilize the tertiary and quaternary levels of protein structure.
b. Peptide bonds form between amino acids in the primary structure of a protein.
c. Hydrophobic interactions between two nonpolar R groups occur in the tertiary and quaternary structures of levels of protein structure.

STUDY CHECK 18.9

What structural level is represented by the grouping of two subunits in insulin?

ANSWER

quaternary structural level

QUESTIONS AND PROBLEMS

18.5 Protein Structure

LEARNING GOAL Identify the levels of structure of a protein.

18.43 Two peptides each contain one molecule of valine and two molecules of serine. Write the three-letter and one-letter abbreviations for their possible primary structures.

18.44 Two peptides each contain one molecule of alanine, one molecule of glycine, and one molecule of isoleucine. Write the three-letter and one-letter abbreviations for their possible primary structures.

18.45 What is the difference in bonding between an α helix and a β-pleated sheet?

18.46 How is the structure of a β-pleated sheet different from that of a triple helix?

18.47 What type of interaction would you expect between the R groups of the following amino acids in the tertiary structural level of a protein?

a. cysteine and cysteine
b. glutamate and arginine
c. serine and aspartate
d. leucine and leucine

d. leucine and leucine

18.48 What type of interaction would you expect between the R groups of the following amino acids in the tertiary structural level of a protein?

a. phenylalanine and isoleucine
b. aspartate and histidine
c. asparagine and tyrosine
d. alanine and proline

18.49 Indicate whether each of the following statements describes the primary, secondary, tertiary, or quaternary protein structure:

a. R groups in amino acids interact to form disulfide bonds or ionic bonds.
b. Peptide bonds join the amino acids in a polypeptide chain.
c. Several polypeptides in a β-pleated sheet are held together by hydrogen bonds between adjacent chains.

18.50 Indicate whether each of the following statements describes the primary, secondary, tertiary, or quaternary protein structure:

a. Hydrogen bonding between amino acid R groups in the same polypeptide gives a coiled shape to the protein.
b. Hydrophilic amino acids move to the polar aqueous environment outside the protein.
c. An active protein contains four subunits.

18.6 Proteins as Enzymes

LEARNING GOAL Describe the role of an enzyme in an enzyme-catalyzed reaction.

An enzyme has a unique three-dimensional shape that recognizes and binds a small group of reacting molecules called substrates. In a catalyzed reaction, an enzyme must first bind to a substrate in a way that favors catalysis.
A typical enzyme is much larger than its substrate. However, within its tertiary structure is a region called the **active site** that binds a substrate or substrates and catalyzes the reaction. This active site is often a small pocket within the larger tertiary structure that closely fits the substrate (see **FIGURE 18.14**).

![FIGURE 18.14](image)

**FIGURE 18.14**  ► On the surface of an enzyme, a small region called the active site binds a substrate and catalyzes a reaction of that substrate.

Why does an enzyme catalyze a reaction of only certain substrates?

**Enzyme-Catalyzed Reaction**

The combination of an enzyme (E) and a substrate (S) within the active site forms an enzyme–substrate (ES) complex that provides an alternative pathway for the reaction with lower activation energy. Within the active site, the amino acid R groups catalyze the reaction to give an enzyme–product (EP) complex. Then the products are released, and the enzyme is available to bind to another substrate molecule.

\[
E + S \rightleftharpoons \text{ES complex} \rightarrow \text{EP complex} \rightarrow E + P
\]

**Models of Enzyme Action**

An early theory of enzyme action, called the **lock-and-key model**, described the active site as having a rigid, nonflexible shape. According to the lock-and-key model, the shape of the active site was analogous to a lock, and its substrate was the key that specifically fit that lock. However, this model was a static one that did not allow for the flexibility of the tertiary shape of an enzyme and the way we now know that the active site can adjust to the shape of a substrate.

In the dynamic model of enzyme action, called the **induced-fit model**, the flexibility of the active site allows it to adapt to the shape of the substrate. At the same time, the shape of the substrate may be modified to better fit the geometry of the active site. As a result, the fit of both the active site and the substrate provides the best alignment for the catalysis of the reaction of the substrate. In the induced-fit model, substrate and enzyme work together to acquire a geometrical arrangement that lowers the activation energy.

In the hydrolysis of the disaccharide lactose by the enzyme lactase, a molecule of lactose binds to the active site of lactase. As the lactose binds to the enzyme, both the active site and the substrate lactose change shape. In this ES complex, the glycosidic bond of lactose is in a position that is favorable for hydrolysis, which is the splitting by water of a large molecule into smaller parts. The R groups on the amino acids in the active site then catalyze the hydrolysis of lactose which produces the monosaccharides glucose and galactose. Because the structures of the products are no longer attracted to the active site, they are released, which allows lactase to react with another lactose molecule (see **FIGURE 18.15**).
In the induced-fit model, a flexible active site and substrate both adjust to provide the best fit for the reaction. Lactose binds to the active site to align the glycosidic bond for hydrolysis. The monosaccharide products are released, and the enzyme binds to another lactose.

**Why does the enzyme-catalyzed hydrolysis of lactose go faster than the hydrolysis of lactose in the chemistry laboratory?**

**QUESTIONS AND PROBLEMS**

18.6 Proteins as Enzymes

**LEARNING GOAL** Describe the role of an enzyme in an enzyme-catalyzed reaction.

18.51 Match the terms, (1) enzyme, (2) enzyme–substrate complex, and (3) substrate, with each of the following:
- a. has a tertiary structure that recognizes the substrate
- b. has a structure that fits the active site of an enzyme
- c. the combination of an enzyme with the substrate

18.52 Match the terms, (1) active site, (2) lock-and-key model, and (3) induced-fit model, with each of the following:
- a. the portion of an enzyme where catalytic activity occurs
- b. an active site that adapts to the shape of a substrate
- c. an active site that has a rigid shape

18.53 a. Write an equation that represents an enzyme-catalyzed reaction.

b. How is the active site different from the whole enzyme structure?

18.54 a. Why does an enzyme speed up the reaction of a substrate?

b. After the products have formed, what happens to the enzyme?

18.7 Nucleic Acids

**LEARNING GOAL** Describe the structure of the nucleic acids in DNA and RNA.

There are two closely related types of nucleic acids: deoxyribonucleic acid (DNA) and ribonucleic acid (RNA). Both are polymers of repeating monomer units known as nucleotides. A DNA molecule may contain several million nucleotides; smaller RNA molecules may contain up to several thousand. Each nucleotide has three components: a base that contains nitrogen, a five-carbon sugar, and a phosphate group (see **Figure 18.16**).

**Bases**

The nitrogen-containing bases in nucleic acids are derivatives of purine or pyrimidine. In DNA, the purine bases with double rings are adenine (A) and guanine (G), and the pyrimidine bases with single rings are cytosine (C) and thymine (T). RNA contains the same bases, except thymine (5-methyluracil) is replaced by uracil (U) (see **Figure 18.17**).
**Ribose and Deoxyribose Sugars**

In RNA, the five-carbon sugar is *ribose*, which gives the letter R in the abbreviation RNA. The atoms in the pentose sugars are numbered with primes (1', 2', 3', 4', and 5') to differentiate them from the atoms in the bases. In DNA, the five-carbon sugar is *deoxyribose*, which is similar to ribose except that there is no hydroxyl group (—OH) on C2'. The *deoxy* prefix means “without oxygen” and provides the D in DNA (see **FIGURE 18.18**).

**Nucleosides and Nucleotides**

A **nucleoside** is produced when a pyrimidine or purine forms a glycosidic bond to C1’ of a sugar, either ribose or deoxyribose. For example, adenine, a purine, and ribose form a nucleoside called adenosine.
Nucleotides are produced when the C5′—OH group of ribose or deoxyribose in a nucleoside forms a phosphate ester. All the nucleotides in RNA and DNA are shown in Figure 18.19.

The name of a nucleoside that contains a purine ends with *osine*, whereas a nucleoside that contains a pyrimidine ends with *idine*. The names of nucleosides of DNA add *deoxy* to the beginning of their names. The corresponding nucleotides in RNA and DNA are named by adding *monophosphate* to the end of the nucleoside name. Although the letters A, G, C, U, and T represent the bases, they are often used in the abbreviations of the respective nucleotides.

Structure of Nucleic Acids

The nucleic acids are polymers of many nucleotides in which the 3′ hydroxyl group of the sugar in one nucleotide bonds to the phosphate group on the 5′ carbon atom in the sugar of the next nucleotide. This connection between the sugars in adjacent nucleotides is referred to as a *phosphodiester linkage*. As more nucleotides are added, a backbone forms that consists of alternating sugar and phosphate groups. The bases, which are attached to each sugar, extend out from the sugar–phosphate backbone.

In any nucleic acid, the sugar at one end has a free 5′ phosphate group, and the sugar at the other end has a free 3′ hydroxyl group. A nucleic acid sequence is read from the free 5′ phosphate end to the free 3′ hydroxyl end using only the letters of the bases. For example, the nucleotide sequence in the section of RNA shown in Figure 18.20 is —A C G U—.
In 1953, James Watson and Francis Crick proposed that DNA was a double helix that consists of two polynucleotide strands winding about each other like a spiral staircase. The sugar–phosphate backbones are analogous to the outside railings, with the bases arranged like steps along the inside.

Complementary Base Pairs
Each of the bases along one polynucleotide strand forms hydrogen bonds to a specific base on the opposite DNA strand. Adenine only bonds to thymine, and guanine only bonds to cytosine (see Figure 18.21). The pairs AT and GC are called complementary base pairs. The specific pairing of the bases occurs because adenine and thymine form two hydrogen bonds, while cytosine and guanine form three hydrogen bonds.

**DNA Double Helix: A Secondary Structure**

In 1953, James Watson and Francis Crick proposed that DNA was a double helix that consists of two polynucleotide strands winding about each other like a spiral staircase. The sugar–phosphate backbones are analogous to the outside railings, with the bases arranged like steps along the inside.

**Complementary Base Pairs**
Each of the bases along one polynucleotide strand forms hydrogen bonds to a specific base on the opposite DNA strand. Adenine only bonds to thymine, and guanine only bonds to cytosine (see Figure 18.21). The pairs AT and GC are called complementary base pairs. The specific pairing of the bases occurs because adenine and thymine form two hydrogen bonds, while cytosine and guanine form three hydrogen bonds.

**Figure 18.20** In the primary structure of an RNA, the nucleotides are linked by 3′,5′ phosphodiester linkages.

What is the abbreviation for the sequence of nucleotides in this RNA section?
SAMPLE PROBLEM 18.10 Complementary Base Pairs

Write the complementary base sequence for the following segment of a strand of DNA:

ACGATCT

TRY IT FIRST

SOLUTION

In the complementary strand of DNA, the complementary base pairs are AT and CG.

Original segment of DNA:  A C G A T C T
Complementary segment:  T G C T A G A

STUDY CHECK 18.10

What sequence of bases is complementary to a DNA sequence of GGTATAACC?

ANSWER

CCAATTTGG
DNA Replication

In DNA replication, the strands in the parent DNA separate, which allows the synthesis of complementary strands of DNA. The replication process begins when an enzyme catalyzes the unwinding of a portion of the double helix by breaking the hydrogen bonds between the complementary bases. These single strands of parent DNA now act as templates for the synthesis of new complementary strands of DNA (see Figure 18.22). Within the nucleus, nucleotides for each base form hydrogen bonds with their complementary bases.

Eventually, the entire double helix of the parent DNA is copied. In each new DNA molecule, one strand of the double helix is from the original DNA and one is a newly synthesized strand. This process produces two new DNAs called daughter DNA that are identical to each other and exact copies of the original parent DNA. In DNA replication, complementary base pairing ensures the correct placement of bases in the daughter DNA strands.
18.8 Protein Synthesis

**LEARNING GOAL** Describe the synthesis of protein from mRNA.

Ribonucleic acid, RNA, which makes up most of the nucleic acid found in the cell, is involved with transmitting the genetic information needed to operate the cell. Similar to DNA, RNA molecules are polymers of nucleotides. However, RNA differs from DNA in several important ways:

1. The sugar in RNA is ribose rather than the deoxyribose found in DNA.
2. In RNA, the base uracil replaces thymine.
3. RNA molecules are single stranded, not double stranded.
4. RNA molecules are much smaller than DNA molecules.

**Types of RNA**

There are three major types of RNA in the cells: *messenger RNA*, *ribosomal RNA*, and *transfer RNA*, which are classified according to their location and function, as shown in **TABLE 18.10**.

<table>
<thead>
<tr>
<th>Type</th>
<th>Abbreviation</th>
<th>Function in the Cell</th>
<th>Percentage of Total RNA</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ribosomal RNA</td>
<td>rRNA</td>
<td>major component of the ribosomes; site of protein synthesis</td>
<td>80</td>
</tr>
<tr>
<td>Messenger RNA</td>
<td>mRNA</td>
<td>carries information for protein synthesis from the DNA to the ribosomes</td>
<td>5</td>
</tr>
<tr>
<td>Transfer RNA</td>
<td>tRNA</td>
<td>brings specific amino acids to the site of protein synthesis</td>
<td>15</td>
</tr>
</tbody>
</table>

In replication, the genetic information in DNA is reproduced by making identical copies of DNA. In **transcription**, the information contained in DNA is transferred to mRNA molecules. In **translation**, the genetic information now present in the mRNA is used to build the sequence of amino acids of the desired protein (see **FIGURE 18.23**).
Transcription begins when the section of DNA to be copied unwinds. One strand of DNA acts as a template as bonds are formed to each complementary base: C is paired with G, T pairs with A, and A pairs with U (not T).

**SAMPLE PROBLEM 18.11 RNA Synthesis**

The sequence of bases in a segment of the DNA template for mRNA is C G A T C A. What corresponding mRNA is produced?

**TRY IT FIRST**

**SOLUTION**

To form the corresponding section of mRNA, each base in the DNA template is paired with its complementary base: G with C, C with G, T with A, and A with U.

<table>
<thead>
<tr>
<th>DNA template:</th>
<th>C G A T C A</th>
</tr>
</thead>
<tbody>
<tr>
<td>Transcription</td>
<td>↓↓↓↓↓↓</td>
</tr>
<tr>
<td>Complementary base sequence in mRNA:</td>
<td>G C U A G U</td>
</tr>
</tbody>
</table>

**STUDY CHECK 18.11**

What is the DNA template that codes for the mRNA segment with the nucleotide sequence G G G U U U A A A A?

**ANSWER**

C C C A A A T T T

**The Genetic Code**

The genetic code consists of a series of three nucleotides (triplet) in mRNA, called a codon. Each codon specifies an amino acid and its sequence in a protein. Early work on protein synthesis showed that repeating triplets of uracil, UUU, produced a polypeptide that contained only phenylalanine. Therefore, a sequence of UUU UUU UUU codes for three phenylalanines.
Codons have now been determined for all 20 amino acids. A total of 64 codons is possible from the triplet combinations of A, G, C, and U (see TABLE 18.11). Three of these, UGA, UAA, and UAG, are stop signals that code for the termination of protein synthesis. All the other codons specify amino acids; one amino acid can have several codons. For example, glycine has four codons: GGU, GGC, GGA, and GGG. The triplet AUG has two roles in protein synthesis. At the beginning of an mRNA, the codon AUG signals the start of protein synthesis. In the middle of a series of codons, AUG codes for the amino acid methionine.

### TABLE 18.11 mRNA Codons: The Genetic Code for Amino Acids

<table>
<thead>
<tr>
<th>First Letter</th>
<th>Second Letter</th>
<th>Third Letter</th>
</tr>
</thead>
<tbody>
<tr>
<td>U</td>
<td>U</td>
<td>U</td>
</tr>
<tr>
<td>U</td>
<td>C</td>
<td>C</td>
</tr>
<tr>
<td>U</td>
<td>A</td>
<td>A</td>
</tr>
<tr>
<td>U</td>
<td>G</td>
<td>G</td>
</tr>
<tr>
<td>C</td>
<td>U</td>
<td>C</td>
</tr>
<tr>
<td>C</td>
<td>C</td>
<td>C</td>
</tr>
<tr>
<td>C</td>
<td>A</td>
<td>A</td>
</tr>
<tr>
<td>C</td>
<td>G</td>
<td>G</td>
</tr>
<tr>
<td>A</td>
<td>U</td>
<td>U</td>
</tr>
<tr>
<td>A</td>
<td>C</td>
<td>C</td>
</tr>
<tr>
<td>A</td>
<td>A</td>
<td>A</td>
</tr>
<tr>
<td>A</td>
<td>G</td>
<td>G</td>
</tr>
<tr>
<td>G</td>
<td>U</td>
<td>U</td>
</tr>
<tr>
<td>G</td>
<td>C</td>
<td>C</td>
</tr>
<tr>
<td>G</td>
<td>A</td>
<td>A</td>
</tr>
<tr>
<td>G</td>
<td>G</td>
<td>G</td>
</tr>
</tbody>
</table>

START\(^a\) codon signals the initiation of a peptide chain. STOP\(^b\) codons signal the end of a peptide chain.

**Protein Synthesis: Translation**

Once the mRNA is synthesized, it migrates out of the nucleus into the cytoplasm to the ribosomes. At the ribosomes, the translation process converts the codons on mRNA into amino acids to make a protein.

Protein synthesis begins when the mRNA combines with a ribosome. There, tRNA molecules, which carry amino acids, align with mRNA, and a peptide bond forms between the amino acids. After the first tRNA detaches from the ribosome, the ribosome shifts to the next codon on the mRNA. Each time the ribosome shifts and the next tRNA aligns with the mRNA, a peptide bond joins the new amino acid to the growing polypeptide chain. After all the amino acids for a particular protein have been linked together by peptide bonds, the ribosome encounters a stop codon. Because there are no tRNAs to complement the termination codon, protein synthesis ends and the completed polypeptide chain is released from the ribosome. Then interactions between the amino acids in the chain form the protein into the three-dimensional structure that makes the polypeptide into a biologically active protein (see FIGURE 18.24).
In the translation process, the mRNA synthesized by transcription attaches to a ribosome, and tRNAs pick up their amino acids, bind to the appropriate codon, and place them in a growing peptide chain.

How is the correct amino acid placed in the peptide chain?
**SAMPLE PROBLEM 18.12  Protein Synthesis: Translation**

Use three-letter and one-letter abbreviations to write the amino acid sequence for the peptide from the mRNA sequence of UCA AAA GCC CUU.

**TRY IT FIRST**

**SOLUTION**

Each of the codons specifies a particular amino acid. Using Table 18.11, we write the codons and the amino acids in the peptide.

<table>
<thead>
<tr>
<th>mRNA codons:</th>
<th>UCA</th>
<th>AAA</th>
<th>GCC</th>
<th>CUU</th>
</tr>
</thead>
<tbody>
<tr>
<td>Amino acid sequence:</td>
<td>Ser—Lys—Ala—Leu, SKAL</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**STUDY CHECK 18.12**

Use three-letter and one-letter abbreviations to write the amino acid sequence for the peptide from the mRNA sequence of GGG AGC AGU GAG GUU.

**ANSWER**

Gly–Ser–Ser–Glu–Val, GSSEV
At the lipid clinic, Susan told Rebecca how to maintain a diet with less beef and chicken, more fish with omega-3 oils, low-fat dairy products, no egg yolks, and no coconut or palm oils. Rebecca maintained her new diet containing lower quantities of fats and increased quantities of fiber for the next year. She also increased her exercise using a treadmill and cycling and lost 35 lb. When Rebecca returned to the lipid clinic to check her progress with Susan, a new set of blood tests indicated that her total cholesterol had dropped from 420 to 390 mg/dL and her LDL had dropped from 275 to 255 mg/dL.

Susan prescribed pravastatin (Pravachol), 80 mg once a day, which was effective. Later, Susan added fenofibrate (TriCor), one 145-mg tablet, each day to the Pravachol. Rebecca understands that her medications and diet and exercise plan is a lifelong process and its impact is to help her live a healthier, longer life.

Because Rebecca’s changes in diet and exercise did not sufficiently lower her cholesterol level, Susan prescribed a medication to help lower blood cholesterol levels. The most common medications used for treating high LDL-cholesterol are the statins lovastatin (Mevacor), pravastatin (Pravachol), simvastatin (Zocor), atorvastatin (Lipitor), and rosuvastatin (Crestor). For some FH patients, statin therapy may be combined with fibrates such as gemfibrozil or fenofibrate that reduce the synthesis of enzymes that break down fats in the blood.

Applications

18.69 If Rebecca’s prescription for Pravachol was 80 mg daily, how many grams of Pravachol did Rebecca consume in one week?

18.70 Five months later, Rebecca added the fibrate TriCor, one 145-mg tablet daily. How many grams of TriCor did Rebecca consume in one week?
18.1 Carbohydrates

**LEARNING GOAL** Classify a carbohydrate as an aldose or a ketose; draw the open-chain and Haworth structures for glucose, galactose, and fructose.

- Carbohydrates are composed of carbon, hydrogen, and oxygen.
- Monosaccharides are polyhydroxy aldehydes (aldoses) or ketones (ketoses).
- Monosaccharides are also classified by their number of carbon atoms: triose, tetrose, pentose, or hexose. Important monosaccharides are glucose, galactose, and fructose.
- The predominant form of monosaccharides is the cyclic form of five or six atoms.
- The Haworth structure forms when an \(-\text{OH}\) group (usually the one on carbon 5 in hexoses) reacts with the carbonyl group of the same molecule.

18.2 Disaccharides and Polysaccharides

**LEARNING GOAL** Describe the monosaccharide units and linkages in disaccharides and polysaccharides.

- Disaccharides are two monosaccharide units joined together by a glycosidic bond.
- In the most common disaccharides maltose, lactose, and sucrose, there is at least one glucose unit.
- Polysaccharides are polymers of monosaccharide units. Amylose is an unbranched polymer of glucose, and amylopectin is a branched polymer of glucose. Glycogen, the storage form of glucose in animals, is similar to amylopectin with more branching. Cellulose is also a polymer of glucose, but in cellulose the glycosidic bonds are \(\beta\) bonds rather than \(\alpha\) bonds.
18.3 Lipids
LEARNING GOAL Draw the condensed or line-angle structural formula for a fatty acid, a triacylglycerol, and the products of hydrogenation or saponification. Identify the steroid nucleus.
- Lipids are nonpolar compounds that are not soluble in water.
- Classes of lipids include waxes, fats, oils, and steroids.
- Fatty acids are long-chain carboxylic acids that may be saturated or unsaturated.
- Triacylglycerols are esters of glycerol with three fatty acids.
- Fats contain more saturated fatty acids and have higher melting points than most vegetable oils.
- The hydrogenation of unsaturated fatty acids converts double bonds to single bonds.
- In saponification, a fat heated with a strong base produces glycerol and the salts of the fatty acids (soap).
- Steroids are lipids containing the steroid nucleus, which is a fused structure of four rings.

18.4 Amino Acids and Proteins
LEARNING GOAL Describe protein functions and draw structures for amino acids and peptides.
- Some proteins are enzymes or hormones, whereas others are important in structure, transport, protection, storage, and muscle contraction.
- A group of 20 amino acids provides the molecular building blocks of proteins.
- Attached to the central (alpha) carbon of each amino acid are an ammonium group, a carboxylate group, and a unique R group.
- Peptides form when an amide bond links the carboxylate group of one amino acid and the ammonium group of a second amino acid.
- Long chains of amino acids that are biologically active are called proteins.

18.5 Protein Structure
LEARNING GOAL Identify the levels of structure of a protein.
- The primary structure of a protein is its sequence of amino acids joined by peptide bonds.
- In the secondary structure, hydrogen bonds between atoms in the peptide bonds produce a characteristic shape such as an alpha helix or a beta-pleated sheet.
- The most abundant protein in the body is collagen, which is composed of fibrils of triple helices that are hydrogen bonded.
- A tertiary structure is stabilized by interactions between R groups of amino acids in one region of the polypeptide chain with R groups in different regions of the protein.
- In a quaternary structure, two or more subunits combine for biological activity.

18.6 Proteins as Enzymes
LEARNING GOAL Describe the role of an enzyme in an enzyme-catalyzed reaction.
- Enzymes are proteins that act as biological catalysts by accelerating the rate of cellular reactions.
- Within the tertiary structure of an enzyme, a small pocket called the active site binds the substrate.
- In the lock-and-key model, an early theory of enzyme action, a substrate precisely fits the shape of the active site.
- In the induced-fit model, both the active site and the substrate undergo changes in their shapes to give the best fit for efficient catalysis. In the enzyme–substrate complex, catalysis takes place when the amino acid R groups in the active site of an enzyme react with a substrate.
- When the products of catalysis are released, the enzyme can bind another substrate molecule.

18.7 Nucleic Acids
LEARNING GOAL Describe the structure of the nucleic acids in DNA and RNA.
- Nucleic acids, such as deoxyribonucleic acid (DNA) and ribonucleic acid (RNA), are polymers of nucleotides.
- A nucleotide is composed of three parts: a base, a pentose sugar, and a phosphate group.
- In DNA, the sugar is deoxyribose and the base can be adenine, thymine, guanine, or cytosine. In RNA, the sugar is ribose and uracil replaces thymine.
- Each nucleic acid has its own unique sequence of bases.
- A DNA molecule consists of two strands of nucleotides that are wound around each other like a spiral staircase.
- The two strands are held together by hydrogen bonds between complementary base pairs AT and GC.

18.8 Protein Synthesis
LEARNING GOAL Describe the synthesis of protein from mRNA.
- The bases in the mRNA are complementary to the DNA, except A in DNA is paired with U in RNA.
- The genetic code consists of a series of codons, which are sequences of three bases that specify the amino acids in a protein.
- Proteins are synthesized at the ribosomes.
- During translation, tRNAs bring the appropriate amino acids to the mRNA at the ribosome and peptide bonds form.
- When the polypeptide is released, it takes on its secondary and tertiary structures and becomes a functional protein in the cell.
KEY TERMS

active site A pocket in a part of the tertiary enzyme structure that binds substrate and catalyzes a reaction.

aldose A monosaccharide that contains an aldehyde group.

amino acid The building block of proteins, consisting of a hydrogen atom, an ammonium group, a carboxylate group, and a unique R group attached to the alpha carbon.

amylopectin A branched-chain polymer of starch composed of glucose units joined by \( \alpha(1\rightarrow4) \)- and \( \alpha(1\rightarrow6) \)-glycosidic bonds.

amylose An unbranched polymer of starch composed of glucose units joined by \( \alpha(1\rightarrow4) \)-glycosidic bonds.

carbohydrate A simple or complex sugar composed of carbon, hydrogen, and oxygen.

cellulose An unbranched polysaccharide composed of glucose units linked by \( \beta(1\rightarrow4) \)-glycosidic bonds that cannot be hydrolyzed by the human digestive system.

cholesterol The most prevalent of the steroid compounds found in cellular membranes.

codon A sequence of three bases in mRNA that specifies a certain amino acid to be placed in a protein. A few codons signal the start or stop of protein synthesis.

complementary base pairs In DNA, adenine is always paired with thymine (A and T or T and A), and guanine is always paired with cytosine (G and C or C and G). In forming RNA, adenine is paired with uracil (A and U).

disaccharide A carbohydrate composed of two monosaccharides joined by a glycosidic bond.

DNA Deoxyribonucleic acid; the genetic material of all cells containing nucleotides with deoxyribose, phosphate, and the four bases: adenine, thymine, guanine, and cytosine.

double helix The helical shape of the double chain of DNA that is like a spiral staircase with a sugar–phosphate backbone on the outside and base pairs like stair steps on the inside.

enzyme A protein that catalyzes a biological reaction.

fat A triacylglycerol that is solid at room temperature and usually comes from animal sources.

fatty acid A long-chain carboxylic acid found in many lipids.

fructose A ketohexose which is combined with glucose in sucrose.

galactose An aldohexose that occurs combined with glucose in lactose.

genetic code The sequence of codons in mRNA that specifies the amino acid order for the synthesis of protein.

glucose An aldohexose found in fruits, vegetables, corn syrup, and honey that is also known as blood sugar and dextrose. The most prevalent monosaccharide in the diet. Most polysaccharides are polymers of glucose.

glycogen A polysaccharide formed in the liver and muscles for the storage of glucose as an energy reserve. It is composed of glucose in a highly branched polymer joined by \( \alpha(1\rightarrow4) \)- and \( \alpha(1\rightarrow6) \)-glycosidic bonds.

induced-fit model A model of enzyme action in which the shape of a substrate and the active site of the enzyme adjust to give an optimal fit.

ketose A monosaccharide that contains a ketone group.

lactose A disaccharide consisting of glucose and galactose found in milk and milk products.

lipids A family of compounds that is nonpolar in nature and not soluble in water; includes fats, oils, waxes, and steroids.

maltose A disaccharide consisting of two glucose units; it is obtained from the hydrolysis of starch.

monosaccharide A polyhydroxy compound that contains an aldehyde or a ketone group.

nucleic acid A large molecule composed of nucleotides; found as a double helix in DNA and as the single strands of RNA.

nucleoside The combination of a pentose sugar and a base.

nucleotide Building block of a nucleic acid consisting of a base, a pentose sugar (ribose or deoxyribose), and a phosphate group.

oil A triacylglycerol that is usually a liquid at room temperature and is obtained from a plant source.

phosphodiester linkage The phosphate link that joins the hydroxyl group in one nucleotide to the phosphate group on the next nucleotide.

polysaccharide A polymer of many monosaccharide units, usually glucose. Polysaccharides differ in the types of glycosidic bonds and the amount of branching in the polymer.

primary structure The specific sequence of the amino acids in a protein.

protein A term used for biologically active polypeptides that have many amino acids linked together by peptide bonds.

quaternary structure A protein structure in which two or more protein subunits form an active protein.

replication The process of duplicating DNA by pairing the bases on each parent strand with their complementary bases.

RNA Ribonucleic acid; a type of nucleic acid that is a single strand of nucleotides containing ribose, phosphate, and the four bases: adenine, cytosine, guanine, and uracil.

saponification The reaction of a fat with a strong base to form glycerol and salts of fatty acids (soaps).

secondary structure The formation of an \( \alpha \)-helix or \( \beta \)-pleated sheet by hydrogen bonds.

steroids Types of lipid composed of a multicyclic ring system.

sucrose A disaccharide composed of glucose and fructose; commonly called table sugar or “sugar.”

tertiary structure The folding of the secondary structure of a protein into a compact structure that is stabilized by the interactions of R groups such as salt bridges and disulfide bonds.

transcription The transfer of genetic information from DNA by the formation of mRNA.

translation The interpretation of the codons in mRNA as amino acids in a peptide.

triacylglycerols A family of lipids composed of three fatty acids bonded through ester bonds to glycerol, a trihydroxy alcohol.
Drawing Haworth Structures (18.1)
- The Haworth structure shows the ring structure of a monosaccharide.
- Groups on the right side of the open-chain structure of the monosaccharide are below the plane of the ring, those on the left are above the plane.
- The new –OH group is the α form if it is below the plane of the ring, or the β form if it is above the plane.

Example: Draw the Haworth structure for β-d-idose.

\[
\begin{align*}
&\text{CH}_2 \equiv \text{O} \\
&\text{HO} \rightarrow \text{C} \rightarrow \text{H} \\
&\text{H} \rightarrow \text{C} \rightarrow \text{OH} \\
&\text{HO} \rightarrow \text{C} \rightarrow \text{H} \\
&\text{H} \rightarrow \text{C} \rightarrow \text{OH} \\
&\text{CH}_2 \text{OH}
\end{align*}
\]

Answer: In the β-form, the new –OH group is above the plane of the ring.

\[
\begin{align*}
&\text{CH}_2 \text{OH} \\
&\text{HO} \rightarrow \text{H} \rightarrow \text{O} \rightarrow \text{OH} \\
&\text{H} \rightarrow \text{OH} \rightarrow \text{H} \\
&\text{OH} \rightarrow \text{H} \rightarrow \text{H}
\end{align*}
\]

Identifying Fatty Acids (18.3)
- Fatty acids are unbranched carboxylic acids that typically contain an even number (12 to 20) of carbon atoms.
- Fatty acids may be saturated, monounsaturated with one double bond, or polyunsaturated with two or more double bonds.

Example: State the number of carbon atoms, saturated or unsaturated, and name of the following:

\[
\begin{align*}
&\text{HO} \rightarrow \text{C} \rightarrow \text{H} \\
&\text{C} \rightarrow \text{O} \rightarrow \text{OH}
\end{align*}
\]

Answer: 16 carbon atoms, saturated, palmitic acid

Drawing Structures for Triacylglycerols (18.3)
- Triacylglycerols are esters of glycerol with three long-chain fatty acids.

Example: Draw the line-angle structural formula for the triacylglycerol containing three molecules of palmitic acid (16:0) and give the name.

\[
\begin{align*}
&\text{CH}_2 \equiv \text{O} \\
&\text{HO} \rightarrow \text{CH} \rightarrow \text{C} \rightarrow \text{OH} \\
&\text{H} \rightarrow \text{C} \rightarrow \text{OH} \\
&\text{CH}_2 \text{OH}
\end{align*}
\]

Glyceryl tripalmitate (tripalmitin)

Drawing the Products for the Hydrogenation and Saponification of a Triacylglycerol (18.3)
- The hydrogenation of unsaturated fatty acids of a triacylglycerol in the presence of a Pt or Ni catalyst converts double bonds to single bonds.
- In saponification, a triacylglycerol heated with a strong base produces glycerol and the salts of the fatty acids (soap).

Example: Identify each of the following as hydrogenation or saponification and state the products:
- The reaction of palm oil with NaOH
- The reaction of corn oil and \( \text{H}_2 \) using a nickel catalyst

Answer: 
- a. The reaction of palm oil with NaOH is saponification, and the products are glycerol and the potassium salts of the fatty acids, which is soap.
- b. In hydrogenation, \( \text{H}_2 \) adds to double bonds in corn oil, which produces a more saturated, and thus more solid, fat.

Drawing the Ionized Form for an Amino Acid (18.4)
- The central α carbon of each amino acid is bonded to an ammonium group (–NH₃⁺), a carboxylate group (–COO⁻), a hydrogen atom, and a unique R group.
- The R group gives an amino acid the property of being nonpolar, polar, acidic, or basic.

Example: Draw the condensed structural formula for cysteine.

\[
\begin{align*}
&\text{SH} \\
&\text{CH}_2 \equiv \text{O} \\
&\text{H}_3\text{N} \rightarrow \text{C} \rightarrow \text{C} \rightarrow \text{O}^{-} \\
&\text{H}
\end{align*}
\]
Identifying the Primary, Secondary, Tertiary, and Quaternary Structures of Proteins (18.5)

- The primary structure of a protein is the sequence of amino acids joined by peptide bonds.
- In the secondary structures of proteins, hydrogen bonds between atoms in the peptide bonds produce an alpha helix or a beta-pleated sheet.
- The tertiary structure of a protein is stabilized by R groups that form hydrogen bonds, disulfide bonds, and salt bridges, and hydrophobic groups that move to the center and hydrophilic R groups that move to the surface.
- In a quaternary structure, two or more subunits are combined for biological activity, held by the same interactions found in tertiary structures.

Example: Identify the following as characteristic of the primary, secondary, tertiary, or quaternary structure of a protein:

a. The R groups of two amino acids interact to form a salt bridge.
b. Eight amino acids form peptide bonds.
c. A polypeptide forms an alpha helix.
d. Two amino acids with hydrophobic R groups move toward the inside of the folded protein.
e. A protein with biological activity contains four polypeptide subunits.

Answer:

a. tertiary, quaternary
b. primary
c. secondary
d. tertiary
e. quaternary

Writing the Complementary DNA Strand (18.7)

- During DNA replication, new DNA strands are made along each of the original DNA strands.
- The new strand of DNA is made by forming hydrogen bonds with the bases in the template strand: A with T and T with A; C with G and G with C.

Example: Write the complementary base sequence for the following DNA segment: A T T C G G T A C

Answer: T A A G C C A T G

Writing the mRNA Segment for a DNA Template (18.8)

- Transcription is the process that produces mRNA from one strand of DNA.
- The bases in the mRNA are complementary to the DNA except that A in DNA is paired with U in RNA.

Example: What is the mRNA produced for the DNA segment C G C A T G T C A ?

Answer: G C G U A C A G U

Writing the Amino Acid for an mRNA Codon (18.8)

- The genetic code consists of a sequence of three bases (codons) that specifies the order for the amino acids in a protein.
- The codon AUG signals the start of transcription and codons UAG, UGA, and UAA signal it to stop.

Example: Use three-letter and one-letter abbreviations to write the amino acid sequence for the peptide from the mRNA sequence C C G U A U G G.

Answer: Pro–Tyr–Gly, PYG

UNDERSTANDING THE CONCEPTS

The chapter sections to review are shown in parentheses at the end of each question.

18.71 Melezitose is a saccharide with the following structure: (18.1, 18.2)

Melezitose

a. Is melezitose a mono-, di-, tri-, or polysaccharide?
b. What ketohexose and aldohexose are present in melezitose?

18.72 What are the disaccharides and polysaccharides present in each of the following? (18.1, 18.2)
18.73 Palm oil has a high concentration of saturated fat, especially the 16-carbon saturated fatty acid, palmitic acid. Draw the line-angle structural formula of palmitic acid. (18.3)

18.74 One of the compounds in jojoba wax used to make candles consists of an 18-carbon saturated fatty acid and a 22-carbon saturated alcohol. Draw the condensed structural formula for jojoba wax. (18.3)

18.75 Identify each of the following as a saturated, monounsaturated, or polyunsaturated fatty acid: (18.3)
   a. \( \text{CH}_3\text{-}(-\text{CH}_2)_6\text{-}(-\text{CH=CH-CH}_2)_2\text{-}(-\text{CH}_2)_6\text{-}\text{COOH} \)
   b. linolenic acid

18.76 Identify each of the following as a saturated, monounsaturated, or polyunsaturated fatty acid: (18.3)
   a. \( \text{CH}_3\text{-}(-\text{CH}_2)_{15}\text{-}\text{COOH} \)
   b. \( \text{CH}_3\text{-}(-\text{CH}_2)_{10}\text{-}\text{CH=CH}_2\text{-}\text{COOH} \)

18.77 Seeds and vegetables are often deficient in one or more essential amino acids. The table below shows which essential amino acids are present in each food. (18.4)

<table>
<thead>
<tr>
<th>Source</th>
<th>Lysine</th>
<th>Tryptophan</th>
<th>Methionine</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oatmeal</td>
<td>No</td>
<td>Yes</td>
<td>Yes</td>
</tr>
<tr>
<td>Rice</td>
<td>No</td>
<td>Yes</td>
<td>Yes</td>
</tr>
<tr>
<td>Garbanzo beans</td>
<td>Yes</td>
<td>No</td>
<td>Yes</td>
</tr>
<tr>
<td>Lima beans</td>
<td>Yes</td>
<td>No</td>
<td>No</td>
</tr>
<tr>
<td>Cornmeal</td>
<td>No</td>
<td>No</td>
<td>Yes</td>
</tr>
</tbody>
</table>

Use the table to decide if each food combination provides the essential amino acids lysine, tryptophan, and methionine.
   a. rice and lima beans
   b. rice and oatmeal
   c. oatmeal and lima beans

18.78 Use the table in problem 18.77 to decide if each food combination provides the essential amino acids lysine, tryptophan, and methionine. (18.4)

18.79 For each of the following pairs of R groups, identify the amino acids and the type of interaction that forms between them: (18.4, 18.5)
   a. \( \text{HO--CH}_2\text{-}(-\text{CH}_2)_4\text{-}\text{O} \) and \( \text{CH}_2\text{-}(-\text{CH}_2)_4\text{-}\text{NH}_2 \)
   b. \( \text{O} \) and \( \text{H}_3\text{N}^+\text{-}(-\text{CH}_2)_4\text{-}\text{O}^- \)

18.80 For each of the following pairs of R groups, identify the amino acids and the type of interaction that forms between them: (18.4, 18.5)
   a. \( \text{CH}_2\text{-}\text{SH} \) and \( \text{HS}\text{-}\text{CH}_2\text{-}\text{COOH} \)
   b. \( \text{CH}_3\text{-}\text{CH}\text{-}\text{CH}_3 \) and \( \text{CH}_3\text{-}\text{COOH} \)

18.81 Answer the following questions for the given section of DNA: (18.7, 18.8)
   a. Complete the bases in the parent and new strands.

Parent Strand:

<table>
<thead>
<tr>
<th>C</th>
<th>G</th>
<th>G</th>
</tr>
</thead>
<tbody>
<tr>
<td>T</td>
<td>T</td>
<td>T</td>
</tr>
<tr>
<td>G</td>
<td>C</td>
<td>T</td>
</tr>
</tbody>
</table>

New Strand:

<table>
<thead>
<tr>
<th>F</th>
<th>T</th>
<th>A</th>
</tr>
</thead>
<tbody>
<tr>
<td>T</td>
<td>G</td>
<td>C</td>
</tr>
</tbody>
</table>

b. Using the new strand as a template, write the mRNA sequence:

<table>
<thead>
<tr>
<th>U</th>
<th>U</th>
<th>A</th>
</tr>
</thead>
<tbody>
<tr>
<td>G</td>
<td>C</td>
<td>U</td>
</tr>
<tr>
<td>C</td>
<td>T</td>
<td>A</td>
</tr>
<tr>
<td>A</td>
<td>T</td>
<td>G</td>
</tr>
</tbody>
</table>

C. Write the 3-letter abbreviations for the amino acids that would go into the peptide from the mRNA you wrote in part b.
### 18.82 Answer the following questions for the given section of DNA:

(18.7, 18.8)

**a.** Complete the bases in the parent and new strands.

Parent Strand: 

| T | T | A | C | C | C |

New Strand: 

| C | T | C | C | T | C |

**b.** Using the new strand as a template, write the mRNA sequence:

| | | | | | |

**c.** Write the 3-letter abbreviations for the amino acids that would go into the peptide from the mRNA you wrote in part **b.**

### 18.83 What are the structural differences in d-glucose and d-galactose? (18.1)

### 18.84 What are the structural differences in d-glucose and d-fructose? (18.1)

### 18.85 Draw the Haworth structure for α-d- and β-d-gulose, whose open-chain structure is shown below. (18.1)

### 18.86 From the compounds shown, select those that match the following:

**a.** an aldohexose  

| C O | H |

**b.** an aldopentose  

| C | H |

**c.** a ketohexose  

### 18.87 Gentiobiose, a carbohydrate found in saffron, contains two glucose molecules linked by a β(1→6)-glycosidic bond. Draw the Haworth structure for α-gentiobiose. (18.1, 18.2)

### 18.88 β-Cellobiose is a disaccharide obtained from the hydrolysis of cellulose. It contains two glucose molecules linked by a β(1→4)-glycosidic bond. Draw the Haworth structure for β-cellobiose. (18.1, 18.2)

### 18.89 Draw the condensed structural formula for Met–Gly–Lys at physiological pH. (18.4)

### 18.90 Draw the condensed structural formula for Thr–Gln–Asp at physiological pH. (18.4)

### 18.91 Identify the base and sugar in each of the following nucleosides: (18.7)

**a.** deoxythymidine  

| C | O | H |

**b.** adenosine  

| C | O | H |

**c.** cytidine  

| C | O | H |

**d.** deoxyguanosine

### 18.92 Identify the base and sugar in each of the following nucleotides: (18.7)

**a.** CMP  

| C | O | H |

**b.** dAMP  

| C | O | H |

**c.** dGMP  

| C | O | H |

**d.** UMP

### 18.93 Write the complementary base sequence for each of the following parent DNA segments: (18.7)

**a.** GTATTGCA  

**b.** GCTAGCGTAA  

**c.** GCTGCACTAGT

### 18.94 Write the complementary base sequence for each of the following parent DNA segments: (18.7)

**a.** AGATCGTAC  

**b.** GTCATACGTAC  

**c.** AAGATGCAGTGT

### 18.95 Match the following statements with rRNA, mRNA, or tRNA:

(18.8)

**a.** carries genetic information from the nucleus to the ribosomes

**b.** found in the ribosome

### 18.96 Match the following statements with rRNA, mRNA, or tRNA:

(18.8)

**a.** acts as a template for protein synthesis

**b.** brings amino acids to the ribosomes for protein synthesis
**Challenge Questions**

The following groups of questions are related to the topics in this chapter. However, they do not all follow the chapter order, and they require you to combine concepts and skills from several sections. These questions will help you increase your critical thinking skills and prepare for your next exam.

18.97 Raffinose is a trisaccharide found in green vegetables such as cabbage, asparagus, and broccoli. It is composed of three different monosaccharides. Identify the monosaccharides in raffinose. (18.1, 18.2)

![Raffinose structure](image)

18.98 Lactulose is a non-absorbable disaccharide which is used in constipation treatment. It is composed of a galactose and a fructose molecule joined by a $\beta(1\rightarrow4)$-glycosidic bond. Draw the Haworth structure for lactulose. (18.1, 18.2)

18.99 Sunflower oil can be used to make margarine. A triacylglycerol in sunflower oil contains two linoleic acids and one oleic acid. (18.3)

![Sunflower oil](image)

18.100 What type of interaction would you expect between the following amino acids in a tertiary structure? (18.4, 18.5)

- a. serine and threonine
- b. cysteine and cysteine
- c. lysine and glutamate

18.101 What are some differences between each of the following pairs? (18.4, 18.5)

- a. secondary and tertiary protein structures
- b. essential and nonessential amino acids
- c. polar and nonpolar amino acids

18.102 What are some differences between each of the following pairs? (18.4, 18.5)

- a. an ionic bond (salt bridge) and a disulfide bond
- b. $\alpha$ helix and $\beta$-pleated sheet
- c. tertiary and quaternary structures of proteins

**Applications**

18.103 Endorphins are polypeptides that reduce pain. Use three-letter abbreviations to write the amino acid sequence for the endorphin leucine enkephalin (leu-enkephalin), which has the following mRNA: (18.4, 18.5, 18.8)

```
AUG UAC GGU GGA UUU CUA UAA
```

18.104 Endorphins are polypeptides that reduce pain. Use three-letter abbreviations to write the amino acid sequence for the endorphin methionine enkephalin (met-enkephalin), which has the following mRNA: (18.4, 18.5, 18.8)

```
AUG UAC GGU GGA UUU AUG UAA
```

18.105 Aspartame, which is used in artificial sweeteners, contains the following dipeptide: (18.4, 18.5)

![Aspartame structure](image)

Aspartame is an artificial sweetener.

- a. What are the amino acids in the dipeptide?
- b. What is the name of the dipeptide in aspartame?
- c. Give the three-letter and one-letter abbreviations for the dipeptide in aspartame.

18.106 The tripeptide shown has a strong attraction for Cu$^{2+}$ and is present in human blood and saliva. (18.4, 18.5)

![Tripeptide structure](image)

- a. What are the amino acids in the tripeptide?
- b. What is the name of this tripeptide?
- c. Give the three-letter and one-letter abbreviations for the tripeptide.

18.107 The following sequence is a portion of a DNA template: (18.4, 18.5, 18.8)

```
GTA CGT AGT AAG
```

- a. Write the corresponding mRNA segment.
- b. Write the three-letter and one-letter abbreviations for the corresponding peptide segment.

18.108 The following sequence is a portion of a DNA template: (18.4, 18.5, 18.8)

```
CCA CTA GCT TAG
```

- a. Write the corresponding mRNA segment.
- b. Write the three-letter and one-letter abbreviations for the corresponding peptide segment.
Answers to Selected Questions and Problems

18.1 Hydroxyl groups are found in all monosaccharides along with a carbonyl on the first or second carbon.

18.3 A ketopentose contains hydroxyl and ketone functional groups and has five carbon atoms.

18.5 In the ring portion of the Haworth structure of glucose, there are five carbon atoms and an oxygen.

18.7 a. α form  b. α form

18.9 a. ketohexose  b. aldopentose

18.11 In galactose, the −OH group on carbon 4 extends to the left. In glucose, this −OH group goes to the right.

18.13 a. one molecule of glucose and one molecule of galactose; β(1→4)-glycosidic bond; β-lactose
   b. two molecules of glucose; α(1→4)-glycosidic bond; α-maltose

18.15 a. disaccharide  b. two molecules of glucose  c. α(1→6)-glycosidic bond  d. α

18.17 a. sucrose  b. lactose  c. maltose  d. lactose

18.19 a. cellulose  b. amylose, amylopectin  c. amylase  d. glycogen

18.21 a. saturated  b. polyunsaturated  c. saturated

18.23

\[
\begin{align*}
\text{CH}_2 &- \text{O} - \text{C} - (\text{CH}_2)_6 - \text{CH}_3 \\
\text{CH} &- \text{O} - \text{C} - (\text{CH}_2)_6 - \text{CH}_3 \\
\text{CH}_2 &- \text{O} - \text{C} - (\text{CH}_2)_6 - \text{CH}_3
\end{align*}
\]

18.25 Safflower oil contains fatty acids with two or three double bonds; olive oil contains a large amount of oleic acid, which has only one (monounsaturated) double bond.

18.27

\[
\begin{align*}
\text{CH}_2 &- \text{O} - \text{C} - (\text{CH}_2)_7 - \text{CH} = \text{CH} - (\text{CH}_2)_3 - \text{CH}_3 \\
\text{CH} &- \text{O} - \text{C} - (\text{CH}_2)_7 - \text{CH} = \text{CH} - (\text{CH}_2)_3 - \text{CH}_3 + 3\text{NaOH} \quad \text{Heat} \\
\text{CH}_2 &- \text{O} - \text{C} - (\text{CH}_2)_7 - \text{CH} = \text{CH} - (\text{CH}_2)_3 - \text{CH}_3
\end{align*}
\]

18.29

\[
\begin{align*}
\text{CH}_2 &- \text{O} - \text{C} - (\text{CH}_2)_{12} - \text{CH}_3 \\
\text{CH} &- \text{O} - \text{C} - (\text{CH}_2)_{12} - \text{CH}_3 + 3\text{NaOH} \quad \text{Heat} \\
\text{CH}_2 &- \text{O} - \text{C} - (\text{CH}_2)_{12} - \text{CH}_3
\end{align*}
\]

18.23

\[
\begin{align*}
\text{CH}_2 &- \text{OH} \\
\text{CH} &- \text{OH} + 3\text{Na}^+ - \text{O} - \text{C} - (\text{CH}_2)_{12} - \text{CH}_3 \\
\text{CH}_2 &- \text{OH}
\end{align*}
\]

18.31

\[
\text{H}_3\text{N} - \text{C} - \text{C} - \text{O}^- \quad \text{H}_3\text{N} - \text{C} - \text{C} - \text{O}^- \quad \text{H}_3\text{N} - \text{C} - \text{O}^- \quad \text{H}_3\text{N} - \text{C} - \text{O}^- \quad \text{H}_3\text{N} - \text{C} - \text{C} - \text{O}^-
\]

18.33 a. \(\text{H}_3\text{N} - \text{C} - \text{C} - \text{O}^-\)  b. \(\text{H}_3\text{N} - \text{C} - \text{C} - \text{O}^-\)

18.35 a. nonpolar, hydrophobic  b. polar, neutral, hydrophilic  c. nonpolar, hydrophobic

18.37 a. alanine  b. valine  c. lysine  d. cysteine

18.39

\[
\begin{align*}
\text{CH}_2 &- \text{O} - \text{C} - (\text{CH}_2)_{14} - \text{CH}_3 \\
\text{CH} &- \text{O} - \text{C} - (\text{CH}_2)_{14} - \text{CH}_3 + 3\text{H}_2 \quad \text{Ni} \\
\text{CH}_2 &- \text{O} - \text{C} - (\text{CH}_2)_{14} - \text{CH}_3
\end{align*}
\]
18.39  a. \[ \text{H}_3\text{N} \text{CH}_3\text{O} \text{CH}_2\text{O} \text{SH} \]  
    Ala–Cys, AC

  b. \[ \text{H}_3\text{N} \text{CH}_3\text{O} \text{CH}_2\text{O} \text{C} \text{O} \]  
    Ser–Phe, SF

  c. \[ \text{H}_3\text{N} \text{CH}_3\text{O} \text{CH}_2\text{O} \text{C} \text{OCH}_3 \text{O} \]  
    Gly–Ala–Val, GA V

18.41  a. \[ \text{NH}_2 \text{CH} \text{C} \text{NH}_2 \]  
    Gly–Ala–Val, GAV

  b. \[ \text{H}_3\text{N} \text{CH}_3\text{O} \text{CH}_2\text{O} \text{C} \text{O} \]  
    RYI

  c. \[ \text{H}_3\text{N} \text{CH}_3\text{O} \text{CH}_2\text{O} \text{C} \text{O} \]  
    VWIS

18.43  Val–Ser–Ser, VSS; Ser–Val–Ser, SVS; or Ser–Ser–Val, SSV

18.45  In the α helix, hydrogen bonds form between the oxygen atom in the C═O group and hydrogen in the N—H group in the next turn of the chain. In the β-pleated sheet, side-by-side hydrogen bonds occur between parallel peptides or across sections of a long polypeptide chain.

18.47  a. disulfide bond  
    b. salt bridge  
    c. hydrogen bond  
    d. hydrophobic interaction

18.49  a. tertiary and quaternary  
    b. primary  
    c. secondary

18.51  a. (1) enzyme  
    b. (3) substrate  
    c. (2) enzyme–substrate complex

18.53  a. \[ \text{E} + \text{S} \xrightarrow{\text{ES}} \text{EP} \xrightarrow{\text{E} + \text{P}} \]  
    b. The active site is a region or pocket within the tertiary structure of an enzyme that accepts the substrate, aligns the substrate for reaction, and catalyzes the reaction.

18.55  a. DNA  
    b. both DNA and RNA

18.57  The two DNA strands are held together by hydrogen bonds between the complementary bases in each strand.

18.59  a. T T T T T T  
    b. C C C C C C  
    c. T C A G G T C C A  
    d. G A C A T A T G C A A

18.61  The DNA strands separate to allow each of the bases to pair with its complementary base, which produces two exact copies of the original DNA.

18.65  a. leucine (Leu)  
    b. serine (Ser)  
    c. glycine (Gly)  
    d. arginine (Arg)

18.67  a. AAA CAC UUG GUU GUG GAC  
    b. Lys–His–Leu–Val–Asp, KHLVD

18.69  0.56 g of Pravachol

18.71  a. Melezitose is a trisaccharide.  
    b. Melezitose contains two glucose molecules and a fructose molecule.

18.73  Palmitic acid

18.75  a. polyunsaturated

18.77  a. yes  
    b. no  
    c. no

18.79  a. asparagine and serine; hydrogen bond  
    b. aspartate and lysine; salt bridge

18.81  a.  
    Parent Strand:  
    New Strand:  
    b.  
    c. Arg-Leu-Ala

18.83  They differ only at carbon 4 where the —OH group in glucose is on the right side and in galactose it is on the left side.
18.85 \[ \alpha - \text{d-Gulose} \quad \beta - \text{d-Gulose} \]

18.87

18.89

18.91 a. thymine and deoxyribose  
        b. adenine and ribose  
        c. cytosine and ribose  
        d. guanine and deoxyribose

18.93 a. CATAACGGT  
        b. CGTAATCGCATT  
        c. CCGACGTTCACTC

18.95 a. mRNA  
        b. rRNA

18.97 galactose, glucose, and fructose

18.99 a.
\[
\begin{align*}
\text{CH}_2O & \text{C}-(\text{CH}_2)_7-\text{CH} \equiv \text{CH}_2 \equiv \text{CH}_2 \equiv \text{CH}-(\text{CH}_2)_4-\text{CH}_3 \\
\text{CH}_2O & \text{C}-(\text{CH}_2)_7-\text{CH} \equiv \text{CH}_2 \equiv \text{CH}-(\text{CH}_2)_4-\text{CH}_3 \\
\text{CH}_2O & \text{C}-(\text{CH}_2)_7-\text{CH} \equiv \text{CH}_2 \equiv \text{CH}-(\text{CH}_2)_4-\text{CH}_3 \\
\end{align*}
\]

b.
\[
\begin{align*}
\text{CH}_2O & \text{C}-(\text{CH}_2)_7-\text{CH} \equiv \text{CH}_2 \equiv \text{CH}-(\text{CH}_2)_4-\text{CH}_3 \\
\text{CH}_2O & \text{C}-(\text{CH}_2)_7-\text{CH} \equiv \text{CH}_2 \equiv \text{CH}-(\text{CH}_2)_4-\text{CH}_3 \\
\text{CH}_2O & \text{C}-(\text{CH}_2)_7-\text{CH} \equiv \text{CH}_2 \equiv \text{CH}-(\text{CH}_2)_4-\text{CH}_3 \\
\end{align*}
\]
18.101 a. The secondary structure of a protein depends on hydrogen bonds to form a helix or a pleated sheet. The tertiary structure is determined by the interaction of R groups, which determine the three-dimensional structure of the protein.

b. Nonessential amino acids are synthesized by the body, but essential amino acids must be supplied by the diet.

c. Polar amino acids have hydrophilic side groups; nonpolar amino acids have hydrophobic side groups.

18.103 START–Tyr–Gly–Gly–Phe–Leu–STOP

18.105 a. aspartate and phenylalanine
   b. aspartylphenylalanine
   c. Asp–Phe, DF

18.107 a. CAU GCA UCA UUG
   b. His–Ala–Ser–Leu, HASL
CL.41 A compound called butylated hydroxytoluene, or BHT, with a molecular formula $\text{C}_{15}\text{H}_{24}\text{O}$, is added to preserve foods such as cereal. As an antioxidant, BHT reacts with oxygen in the cereal container, which protects the food from spoilage.

\[ \text{BHT} \]

CI.42 Olive oil contains a high percentage of glyceryl trioleate (triolein). (7.2, 11.7, 12.6, 18.3)

\[ \text{Glycerin} \]

CI.43 A sink drain can become clogged with solid fat such as glyceryl tristearate (tristearin). (8.2, 12.6, 18.3)

\[ \text{Terephthalic acid} \]

\[ \text{Ethylene glycol} \]

CL.44 In response to signals from the nervous system, the hypothalamus secretes a polypeptide hormone known as gonadotropin-releasing factor (GnRF), which stimulates the pituitary gland to release other hormones into the bloodstream.

\[ \text{Gonadotropin-releasing factor (GnRF)} \]

CI.45 The plastic known as PETE (polyethylene terephthalate) is a polymer of terephthalic acid and ethylene glycol. PETE is used to make plastic soft drink bottles and containers for salad dressings, shampoos, and dishwashing liquids. Today, PETE is the most widely recycled of all the plastics; in a single year, $1.5 \times 10^9$ lb of PETE are recycled. After PETE is separated from other plastics, it can be used in polyester fabric, fill for sleeping bags, door mats, and tennis ball containers. The density of PETE is 1.38 g/mL. (2.5, 2.6, 17.6)
a. Draw the condensed structural formula for the ester formed from one molecule of terephthalic acid and one molecule of ethylene glycol.

b. Draw the condensed structural formula for the product formed when a second molecule of ethylene glycol reacts with the ester you drew for the answer in part a.

c. How many kilograms of PETE are recycled in one year?

d. What volume, in liters, of PETE is recycled in one year?

e. Suppose a landfill with an area of a football field and a depth of 5.0 m holds $2.7 \times 10^7$ L of recycled PETE. If all of the PETE that is recycled in a year were placed in landfills, how many would it fill?

Cl.46 Thalassemia is an inherited genetic mutation that limits the production of hemoglobin. If less hemoglobin is produced, there is a shortage of red blood cells (anemia). As a result, the body does not have sufficient amounts of oxygen. In one form of thalassemia, thymine (T) is deleted from section 91 (bold) in the following segment of normal DNA: (18.4, 18.5, 18.8)

```
89 90 91 92 93 94
AGT CAG CTG CAC TGT GAC A...
```

a. Write the complementary (template) strand for this normal DNA segment.

b. Write the mRNA sequence using the (template) strand in part a.

c. What amino acids are placed in the beta chain using the portion of mRNA in part b?

d. What is the order of nucleotides after T is deleted?

e. Write the template strand for the mutated DNA segment.

f. Write the mRNA sequence from the mutated DNA segment using the template strand in part e.

g. What amino acids are placed in the beta chain by the mutated DNA segment?

h. How might the properties of this segment of the beta chain be different from the properties of the normal protein?

i. How might the level of structure in hemoglobin be affected if beta chains are not produced?

---

ANSWERS

Cl.41 a. The $\text{–OH}$ group in BHT is bonded to a carbon atom in an aromatic ring, which means BHT has a phenol functional group.

b. BHT is referred to as an “antioxidant” because it reacts with oxygen in the food container, rather than the food, thus preventing or retarding spoilage of the food.

c. 21 mg of BHT

Cl.43 a. Adding NaOH will saponify the glycercyl tristearate (fat), breaking it up into fatty acid salts and glycerol, which are soluble and will wash down the drain.

b. 

```
\[
\begin{array}{c}
\text{CH}_2-O-C-(\text{CH}_2)_{16}-\text{CH}_3 \\
\text{CH}-O-C-(\text{CH}_2)_{16}-\text{CH}_3 + 3\text{NaOH} \\
\text{CH}_2-O-C-(\text{CH}_2)_{16}-\text{CH}_3 \\
\text{CH}_2-OH \\
\text{CH}-\text{OH} + 3\text{Na}^+ -O-C-(\text{CH}_2)_{16}-\text{CH}_3 \\
\text{CH}_2-OH
\end{array}
\]
```

c. 67.3 mL of a 0.500 M NaOH solution

Cl.45 a. 

```
\[
\text{HO} \quad -C-\text{O} \quad -C-\text{O} \quad \text{CH} \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{OH}
\]
```

b. 

```
\[
\text{HO} \quad -\text{CH}_2 \quad \text{CH}_2 \quad \text{O} \quad \text{C} \quad \text{C} \quad \text{O} \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{OH}
\]
```

c. $6.8 \times 10^8$ kg of PETE

d. $4.9 \times 10^8$ L of PETE

e. 30 landfills
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## CHAPTER 1

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## CHAPTER 2

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Aldehyde
An organic compound that contains a carbonyl group (C=O) bonded to at least one hydrogen atom, 600r, 612–616
naming, 613–614

Aldohexose, 644

Aldose
A monosaccharide that contains an aldehyde group, 643

Alkaloid, 625

Alkaline earth metal
An element in Group 1A (1), 624

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Alkenes
A type of hydrocarbon in which the carbon atoms are connected only by single bonds, 591
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Alkaline earth metal
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Aliphatic compound
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Amide
An organic compound in which the hydroxyl group of a carboxylic acid is replaced by a nitrogen atom, 600r, 626–627

Ammonium ion, 210r

Amphoteric
Substances that act as either an acid or a base in water, 478

Amylase, 652

Amylopectin
A branched-chain polymer of starch composed of glucose units joined by α(1→4)- and α(1→6)-glycosidic bonds, 651

Amylose
An unbranched polymer of starch composed of glucose units joined by α(1→4)-glycosidic bonds, 538, 651

Amphibole
A secondary level of structure, in which hydrogen bonds connect the N—H of one peptide bond with the C=O of a peptide bond farther down the chain to form a coiled or corkscrew structure, 668, 669

Amphiphile
A polar molecule that can form a micelle or a monolayer at an interface between two immiscible phases, 348, 353

Amphoteric
Substances that act as either an acid or a base in water, 478

Anesthetic
A type of drug that is used to block or reduce the transmission of pain signals from the body to the brain, 535–536, 538–539

Amphoteric
Substances that act as either an acid or a base in water, 478

Anion
A negatively charged ion such as Cl⁻, O₂⁻, or SO₄²⁻, 197

Anode
The electrode where oxidation takes place, 535–536, 538–539

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Aromatic compound
A compound that contains the ring structure of benzene, 600r, 606–608
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Atmosphere (atm) A unit equal to the pressure exerted by a column of mercury 760 mm high, 355, 357, 358r
Atmospheric pressure The pressure exerted by the atmosphere, 355, 379
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Atom The smallest particle of an element, 142–145
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Atomic energy, 575
Atomic mass The weighted average mass of all the naturally occurring isotopes of an element, 148–152
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Atomic mass unit (amu) A small mass unit used to describe the mass of extremely small particles such as atoms and subatomic particles; 1 amu is equal to one-twelfth the mass of $^{12}$C atom, 144, 145r

Atomic number A number that is equal to the number of protons in an atom, 146–148, 147r
Atomic radius, 183–184

Atomic size The distance between the outermost electrons and the nucleus, 183–184, 186r

Atomic spectrum A series of lines specific for each element produced by photons emitted by electrons dropping to lower energy levels, 166

Atomic symbol An abbreviation used to indicate the mass number and atomic number of an isotope, 148, 149r
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Avogadro’s law A gas law stating that the volume of a gas is directly related to the number of moles of gas when pressure and temperature do not change, 370–372

Avogadro’s number The number of items in a mole, equal to $6.022 \times 10^23$, 225–227, 226r

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Background radiation, 566, 566r
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Balanced equation The final form of a chemical equation that shows the same number of atoms of each element in the reactants and products, 257–264
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ethyne (acetylene), 599
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Base (chemical) A substance that dissolves in water and produces hydroxide ions (OH-) according to the Arrhenius theory. All bases are hydrogen ion acceptors, according to the Brønsted–Lowry theory, 473–517
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Base dissociation expression $K_a$. The product of the ions from the dissociation of a weak base divided by the concentration of the weak base, 485
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Be (Beta) A unit of activity of a radioactive sample equal to one disintegration per second, 565, 566r
Beeswax, 658, 658r
Belladonna, 625
Bends, scuba diving and, 382

Bent The shape of a molecule with two bonded atoms and one lone pair or two lone pairs, 320–322, 321r
Benzamide, 627

Benzene A ring of six carbon atoms each of which is attached to one hydrogen atom, ($\text{C}_6\text{H}_6$), 606 in aromatics, 606–608
Benzonic acid, 617
Beta decay, 559–561

Beta particle A particle identical to an electron, symbol $\beta^-$ or $\beta$, that forms in the nucleus when a neutron changes to a proton and an electron, 555, 555r
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Blood sugar. See Glucose
Body, human. See Human body
Body fat, 76
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Boiling The formation of bubbles of gas throughout a liquid, 333, 334

Boiling point (bp) The temperature at which liquid changes to gas (boils) and gas changes to liquid (condenses), 333
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Boyle’s law A gas law stating that the pressure of a gas is inversely related to the volume when temperature and moles of the gas do not change, 359–362, 368r

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Brønsted–Lowry acids and bases An acid is a hydrogen ion donor; a base is a hydrogen ion acceptor, 477–479

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Buffer solution A solution of a weak acid and its conjugate base or a weak base and its conjugate acid that maintains the pH by neutralizing added acid or base, 502–507
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Carbon monoxide, 268
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Carbohydrate An organic compound that contains the carboxyl group (—COOH), 616
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registered nurse, 57
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Casein, 663t
Catalyst A substance that increases the rate of reaction by lowering the activation energy, 443
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Catalytic converters, 444
Cathode The electrode where reduction takes place, 535, 538–539
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Cathodic protection, 540
Cation A positively charged ion, such as Na⁺, Mg²⁺, Al³⁺, or NH₄⁺, 185, 197
Cellulose An unbranched polysaccharide composed of glucose units linked by β(1→4)-glycosidic bonds that cannot be hydrolyzed by the human digestive system, 643, 651–652
Celsius (°C) temperature scale A temperature scale on which water has a freezing point of 0 °C and a boiling point of 100 °C, 44, 60, 107–114
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Centimeter (cm) A unit of length in the metric system; there are 2.54 cm in 1 in., 59
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Cesium-131, 572t, 574
Chain reaction A fission reaction that will continue once it has been initiated by a high-energy neutron bombarding a heavy nucleus such as uranium-235, 576, 577
Chain termination, 682
Change of state The transformation of one state of matter to another, for example, solid to liquid, liquid to solid, liquid to gas, 103–106, 331–339
Changes of energy levels, 167–168
Charles, Jacques, 362–363
Charles’s law A gas law stating that the volume of a gas is directly related to the Kelvin temperature when pressure and moles of the gas do not change, 362–364, 368t
Chemical A substance that has the same composition and properties wherever it is found, 32
in kitchen, 33
in toothpaste, 33t
Chemical bond, 213–217
Chemical change The transformation of one substance into a new substance that has a different composition and new physical and chemical properties, 105, 105t, 255
Chemical compound. See Compound
Chemical equation A shorthand way to represent a chemical reaction using chemical formulas to indicate the reactants and products and coefficients to show reacting ratios, 256
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Chemical equilibrium The point at which the rate of forward and reverse reactions are equal so that no further change in concentrations of reactants and products takes place, 445–448
See also Equilibrium
Chemical formula The group of symbols and subscripts that represent the atoms or ions in a compound, 202–203
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Chemical oceanographer, 439
Chemical properties The properties that indicate the ability of a substance to change into a new substance, 105, 105t
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Chemical symbol An abbreviation that represents the name of an element, 133–134
Chemistry The study of the composition, structure, properties, and reactions of matter, 32–34
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Cholesterol The most prevalent of the steroid compounds found in cellular membranes, 662
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Chinchona tree, 625
Citric acid cycle, 618
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Codon A sequence of three bases in mRNA that specifies a certain amino acid to be placed in a protein. A few codons signal the start or stop of protein synthesis, 681–682
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Coefficients Whole numbers placed in front of the formulas to balance the number of atoms or moles of atoms of each element on both sides of an equation, 257, 462–463
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Collagen 663r, 668
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Collision theory A model for a chemical reaction stating that molecules must collide with sufficient energy and proper orientation to form products, 440, 441
Colloid A mixture having particles that are moderately large. Colloids pass through filters but cannot pass through semipermeable membranes, 419, 419t
comparison to solution and suspension, 420r
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Combination reaction A chemical reaction in which reactants combine to form a single product, 264–265, 267t
Combined gas law A relationship that combines several gas laws relating pressure, volume, and temperature when the amount of gas does not change, 368–369, 368r
Combustion reaction A chemical reaction in which a fuel containing carbon and hydrogen reacts with oxygen to produce CO2, H2O, and energy, 267–268, 267t
Compact fluorescent light (CFL), 167–168
Complementary base pairs In DNA, adenine is always paired with thymine (A and T or T and A), and guanine is always paired with cytosine (G and C or C and G). In forming RNA, adenine is paired with uracil (A and U), 677–678
Complete proteins, 665
Compound A pure substance consisting of two or more elements, with a definite composition, that can be broken down into simpler substances only by chemical methods, 100, 195–223
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Conversion factor A ratio in which the numerator and denominator are quantities from an equality or given relationship, 73–78
concentrations used as, 405, 410r
density as, 88
Metric–U.S. system, 74–75
molarity, 408–409, 410r
molar volume, 371
percentage, ppm, and ppb, 76–77
percent concentration, 410r
with powers, 75
problem solving with, 78–84
stated within a problem, 75–76
Cooling curve A diagram that illustrates temperature changes and changes of states for a substance as heat is removed, 334–336
Coronary heart disease, 553, 662
Corrosion, 539–540
Covalent bond A sharing of valence electrons by atoms, 213–217, 311
forming, 196, 213
nonpolar, 324
polar, 324
resonance structures, 316–318
Crenation The shrinking of a cell due to water leaving the cell when the cell is placed in a hypertonic solution, 426, 427
Crick, Francis, 677
Crude oil, 598
C-terminal amino acid, 666
C-terminus, 666
CT scan. See Computed tomography (CT)
Cubic centimeter (cm3, cc) The volume of a cube that has 1-cm sides; 1 cm3 is equal to 1 mL, 72
Cubic meter (m3) The volume of a cube that has sides that measure 1 m in length, 59
Curie (Ci) A unit of activity of a radioactive sample equal to 3.7 × 1010 disintegrations/s, 564–565, 566r
Cysteine, 664r
Cytosine (C), 674, 675
D
Dalton, John, 142, 145
Dalton’s law A gas law stating that the total pressure exerted by a mixture of gases in a container is the sum of the partial pressures that each gas would exert alone, 378–382
Dating, with half-lives, 570
d block The 10 elements in Groups 3B (13) to 2B (12) in which electrons fill the five d orbitals, 177
Dead Sea Scrolls, 570
Decay curve A diagram of the decay of a radioactive element, 568
Deci (d), 69, 70r
Deciliter (dL), 71
Decimal numbers, 240–241
Decimal point, 47
power of ten, 47–48, 48r
Decomposition reaction A reaction in which a single reactant splits into two or more simpler substances, 265, 267t
Decompression, 382
Deforestation, 114
Denominator, 73
**Density** The relationship of the mass of an object to its volume expressed as grams per cubic centimeter (g/cm³), grams per milliliter (g/mL), or grams per liter (g/L), 84–90
bone density, 87
calculation of, 84–86
of common substances, 85t
problem solving, 88
of solids, 85t, 86–87
volume displacement to calculate, 86–87

**Dentist**, 521
**Deoxyribonucleic acid. See DNA**
**Deoxyribose**, 675
**Deoxyxymidine monophosphate, 676**

**Deposition** The change of a gas directly to a solid; the reverse of sublimation, 338
**Dextrins**, 652
**Dextrose. See Glucose**

**Diabetes, type 2**, 644
See **Dextrose**.

**Diamond**, 135

**Dialyzing membrane**, 427
**Dialysis nurse**, 390

**Dialysis** A process in which water and small solute particles pass through a semipermeable membrane, 390, 427–429
Dialysis nurse, 390
Dialyzing membrane, 427
Diamond, 135
Diaphragm, 361
**Diastolic blood pressure**, 357
**Diatomic molecule**, 311, 311

**Dipole** The separation of positive and negative charge in a polar bond indicated by an arrow that is drawn from the more positive atom to the more negative atom, 324, 325
**Dipole–dipole attractions** Attractive forces between oppositely charged ends of polar molecules, 329, 330t
Diprotic acids, 481–482

**Direct relationship** A relationship in which two properties increase or decrease together, 45, 363

**Disaccharide** A carbohydrate composed of two monosaccharides joined by a glycosidic bond, 643, 647–654
sweetness of, 650t

**Dispersion forces** Weak dipole bonding that results from a momentary polarization of nonpolar molecules, 329, 329t

**Dissociation** The separation of an acid or base into ions in water, 395, 480
constant for acid and bases, 484–487, 485t
Dissociation of water, 487–490, 488t
Disulfide bond, 668
Diver. See **Scuba diver**
Division, 41, 66–67

**DNA** Deoxyribonucleic acid; the genetic material of all cells containing nucleotides with deoxyribose, phosphate, and the four bases: adenine, thymine, guanine, and cytosine, 674–680
base pairs in, 677, 678
bases in, 674–675
components of, 674–680
daughter DNAs, 679
deoxyribose in, 675
double helix, 677–679
nucleotides and nucleosides in, 675–676
primary structure, 676–677
replication, 679
synthesis, 679
d orbitals, 170
Dosage, 82
Dosimeter, 566
**Double bond** A sharing of two pairs of electrons by two atoms, 314–315
**Double helix** The helical shape of the double chain of DNA that is like a spiral staircase with a sugar–phosphate backbone on the outside and base pairs like stair steps on the inside, 677–679, 678

**Double replacement reaction** A reaction in which the positive ions in the reacting compounds exchange places, 266, 267t

**Dry-cell batteries**, 538–539

**E**
**Einstein, Albert**, 576

**Electrical charge** in atom, 143
static electricity, 143

**Electrical energy** batteries for, 538–539
oxidation–reduction reactions, 533–542
voltaic cell for, 535–537
Electrode, 535

**Electrolysis** The use of electrical energy to run a nonspontaneous oxidation–reduction reaction in an electrolytic cell, 542

**Electrolyte** A substance that produces ions when dissolved in water; its solution conducts electricity, 395–396
in body fluids, 398–399

**Electrolytic cell** A cell in which electrical energy is used to make a nonspontaneous oxidation–reduction reaction happen, 542

**Electromagnetic radiation** Forms of energy such as visible light, microwaves, radio waves, infrared, ultraviolet light, and X-rays that travel as waves at the speed of light, 162–165

**Electromagnetic spectrum** The arrangement of types of radiation from long wavelengths to short wavelengths, 163–164

**Electron** A negatively charged subatomic particle having a minute mass that is usually ignored in mass calculations; its symbol is \( e^− \), 143, 145t
energy levels, 165–168
energy sublevels, 168–169, 171t
high-energy (beta particle), 555
orbitals, 168–172
sharing, 213–217
transfer of, 196–201
valence, 181–182, 182t

**Electron configuration** A list of the number of electrons in each sublevel within an atom, arranged by increasing energy, 172–176
abbreviated, 174
Period 1, 173–174
Period 2, 174–175
Period 3, 175–176
Period 4, 178–180
periodic table and, 177–181
writing using sublevel blocks, 177–178

**Electronegativity** The relative ability of an element to attract electrons in a bond, 323–326, 325t

**Electroplating**, 543

**Element** A pure substance containing only one type of matter, which cannot be broken down by chemical methods, 100, 133–136
atom, 142–145
atomic mass, 144–145
covalent bond, 213–217
diatomic molecule, 311, 311t
essential to health, 140, 141t
ionic charge, 196, 198t, 199
Latin names in clinical usage, 133–134
molar mass, 229–231
names, 133t, 134t
periodic table of, 136–142
representative, 136–137
symbols of, 133t, 134t
trace, 140, 141t
transition, 137
See also **Periodic table**

**Empirical formula** The simplest or smallest whole-number ratio of the atoms in a formula, 238–242, 238t
calculating, 238–241
decimal to whole numbers conversion, 240
mass percent composition used to calculate, 239
molecular formula and, 243
multipliers used to calculate, 240

**Endothermic reaction** A reaction wherein the energy of the products is higher than that of the reactants, 296–297
in cold pack, 298
equilibrium shifting from temperature change, 459–460, 460t

**Endpoint** The point at which an indicator changes color. For the indicator phenolphthalein, the color change occurs when the number of moles of OH⁻ is equal to the number of moles of H₂O⁺ in the sample, 500

**Energy** The ability to do work, 111–114
activation, 441
adult requirements, 121t
atomic, 575
caloric value of foods, 119–121, 120t
changes in chemical reactions, 295–300
collision theory, 440–441
fat as a source of, 120t
heat and, 112
ionization, 185
kinetic, 111–112
levels, atomic, 165–168
nuclear fission, 575–576
nuclear fusion, 576
nuclear power plant, 578–579
Energy (Continued)
nutrition and, 119–122, 120r
potential, 111–112
specific heat, 115–118, 115r
units of, 112–113
Energy level A group of electrons with similar energy, 166
Energy value The kilocalories (or kilojoules) obtained per gram of the food types: carbohydrate, fat, and protein, 119, 120r
Enthalpy change, 295
Environmental Protection Agency (EPA)mercury in fish, 136
radon, 559
Environmental scientist, 276
Enzyme A protein that catalyzes a biological reaction, 663r, 672–674
action of, 673–674
active site, 673
as catalyst, 673
catalyzed reactions, 673
induced-fit model, 673, 674
lock-and-key model, 673
proteins as, 672–674
Enzyme-catalyzed reaction, 673
Enzyme–product (EP) complex, 673
Enzyme–substrate (ES) complex, 673
Enzyme-catalyzed reaction, 673
Enzyme–product (EP) complex, 673
Enzyme–substrate (ES) complex, 673
Epinephrine, 461
Equality A relationship between two units that measure the same quantity, 71, 91
common, 74r
conversion factors for, 73–78
length, 70–71
mass, 72
prefixes and, 70r
volume, 71–72
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heterogenous and homogenous, 451
oxygen–hemoglobin, 459
rate of reaction, 440–445, 443r
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state of reactants and products at, 445–448
temperature change and, 442, 443r, 459–461, 460r
volume (pressure) change and, 458, 460r
Equilibrium constant, \( K \)
The numerical value obtained by substituting the equilibrium concentrations of the components into the equilibrium expression, 448–449
calculating, 449–450, 450r
large, 452, 453, 454r
small, 453, 454r
using, 452–455
Equilibrium expression The ratio of the concentrations of products to the concentrations of reactants, with each component raised to an exponent equal to the coefficient of that compound in the balanced chemical equation, 448–449
Equivalent dose The measure of biological damage from an absorbed dose that has been adjusted for the type of radiation, 565
Erythrose, 242r
Essential amino acid, 665–666
Ester An organic compound in which the —H of a carbon group is replaced by a carbon atom, 600r, 619, 622
bond, 658
formation of, 619
fruit and flavoring, 621, 621r
naming of, 620
in plants, 621–622
Esterification, 619–620
Ethanamine, 626–627
Ethane, 591, 592r
Ethanediol (ethylene glycol), 611
Ethanoic acid, 617r, 619–620, 626
Ethanol (ethyl alcohol)
in hand sanitizers, 610
heat of fusion for, 335
heat of vaporization for, 335
value, 82r
See also Alcohol
Ethe (ethylene), 599, 604r
Ether An organic compound in which the hydrogen atom is bonded to two carbon atoms, 608, 609, 610–612
and alcohol, 608–612
Ethyl alcohol, See Ethanol
Ethyl butanoate, 621r
Ethylene (ethene), 599, 604r
Ethylene glycol, 611
Ethyle heptanoate, 622
Ethine (acetylene), 599, 600r
Eugenol, 611
Evaporation The formation of a gas (vapor) by the escape of high-energy molecules from the surface of a liquid, 333, 334
Exact number A number obtained by counting or by definition, 63–64, 63r
Excess reactant, 289
Exercise physiologist, 254
Exhalation, 361
Exothermic reaction A reaction wherein the energy of the products is lower than that of the reactants, 295–296
exhaling, 361
Exothermic reaction A reaction wherein the energy of the products is lower than that of the reactants, 295–296
f-block The 14 elements in the rows at the bottom of the periodic table in which electrons fill the seven 4f and 5f orbitals, 177
Fermi (f), 70r
Ferritin, 663r
Fertilizers, 132, 237
Fibril, collagen, 668
Filtration, 102
Firefighter/emergency medical technician, 588
Fish, mercury level in, 135–136
Fission A process in which large nuclei are split into smaller pieces, releasing large amounts of energy, 575, 577
Flammability inorganic compound, 590r
organic compound, 590r
Food energy values of, 119–121
irradiating, 565
mercury content, 135
nutrition labeling, 120
water percentage, 393
Force attractive, 329, 330r
dispersion, 329, 330r
intermolecular, 328–331
Forensic scientist, 31
Formaldehyde, 271, 319, 610, 613, 632r
Formic acid, 271, 484, 617r
equilibrium equation, 485
Formula The group of symbols and subscripts that represent the atoms or ions in a compound, 202–204
See also Condensed structural formula; Expanded structural formula
Formula unit The group of ions represented by the formula of an ionic compound, 203, 226
Forward reaction, 445, 446
Fossil fuel, 112, 114
Freeze-dried foods, 338
Freezer burn, 338
Freezing A change of state from liquid to solid, 332
**Freezing point (fp)** The temperature at which a liquid changes to a solid (freezes) and a solid changes to a liquid (melts), 332

Celsius (°C) scale, 107, 108r

Fahrenheit (°F) scale, 107, 108r

Kelvin (K) scale, 108, 109r

lowering of, 422

solutes used to lower, 420–421, 424r

**Frequency** The number of times the crests of a wave pass a point in 1 s, 162

**Fructose** A ketohexose which is combined with glucose in sucrose, 643

Fructose, 643

Fruits

esters in, 621r

ripening, 600

Fuel cells, 541–542

Fumaric acid, 619

**Functional group** A group of atoms that determines the physical and chemical properties and naming of a class of organic compounds, 599, 600r

alcohol, 608, 609

aldehyde, 612

alkene, 599

alkyne, 599

amide, 626

amine, 623

carboxylic acid, 616

ester, 619

ether, 608, 609

ketone, 612

**Fusion** A reaction in which large amounts of energy are released when small nuclei combine to form larger nuclei, 576, 577

**G**

**Galactose** An aldohexose that occurs combined with glucose in lactose, 643

Gallium-67, 572r

Gallium-68, 572r

Galvanization, 540

Gamma emission, 562

**Gamma ray** High-energy radiation, symbol a, emitted by an unstable nucleus, 555, 555r

killing bacteria in food with, 565

protecting from, 556, 557r

Gas A state of matter that does not have a definite shape or volume, 104, 104r, 353–389

blood gases, 382

collection of, over water, 380–381

density of common gases, 85r

kinetic molecular theory of, 354–355

laws, 359–382, 368r

molar mass of, 375–376

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partial pressures (Dalton’s law), 378–381

pressure, 355, 356r, 357–358, 400

properties of, 104r, 354–359, 356r, 371

as solution, 392r

STP, 371–372

volume and moles (Avogadro’s law), 370–372

See also Gas laws

Gas laws

combined gas law, 368–369

ideal gas law, 373–376

partial pressures (Dalton’s law), 378–382

pressure and volume (Boyle’s law), 359–362

summarized, 368r

temperature and pressure (Gay-Lussac’s law), 365–367

temperature and volume (Charles’ law), 362–364

temperature (Kelvin) in calculations of, 363, 365

volume and moles (Avogadro’s law), 370–372

**Gay-Lussac’s law** A gas law stating that the pressure of a gas is directly related to the Kelvin temperature when the number of moles of a gas and its volume do not change, 365–367, 368r

Geiger counter, 564

**Genetic code** The sequence of codons in mRNA that specifies the amino acid order for the synthesis of protein, 681–682

mRNA codons for amino acids, 681–682

Gene information. See DNA; RNA

**Giga (G)**, 70r

Glomerulus, 427

**Glucose** An aldohexose found in fruits, vegetables, corn syrup, and honey that is also known as blood sugar and dextrose. The most prevalent monosaccharide in the diet. Most polysaccharides are polymers of glucose, 642–643

blood, 643

molecular and empirical formulas, 238t

Glutamate, 664r

Glutamine, 664r

Glycine, 664r, 666

mRNA codons, 682

**Glycosidic bond** A bond formed between monosaccharides in a polysaccharide molecule, 647

**Glycine** A protein building block, 664

Hemoglobin, 663

Heart attack, 662

Heat The energy associated with the motion of particles in a substance, 112

calculating in reaction, 297–298

calculations using specific heat, 115–116

calculating in reaction, 297–298

calculations using specific heat, 115–116

of condensation, 333–334

specific, 115–118, 115r

units of, 112

**Heat equation** A relationship that calculates heat (q) given the mass, specific heat, and temperature change for a substance, 116–118

**Heating curve** A diagram that illustrates the temperature changes and changes of state of a substance as it is heated, 334–335

**Heat of fusion** The energy required to melt exactly 1 g of a substance at its melting point. For water, 334 J is needed to melt 1 g of ice; 334 J is released when 1 g of water freezes, 332–333

**Heat of reaction** The heat (symbol ΔH) absorbed or released when a reaction takes place at constant pressure, 295

**Heat of vaporization** The energy required to vaporize 1 g of a substance at its boiling point. For water, 2260 J is needed to vaporize exactly 1 g of water; 1 g of steam gives off 2260 J when it condenses, 333–334

Heliox, 103

Helix
double, 677–679

triple, 668, 669

Hemlock, 625

**Hemodialysis** A cleansing of the blood by an artificial kidney using the principle of dialysis, 427–428

Hemoglobin, 663r
carbon monoxide binding to, 268
carrying oxygen, 268, 663, 663r

function of, 663

Guanosine monophosphate, 676

Guide to Problem Solving (GPS), 79

**H**

Hair

mercury testing and, 135

number and width of, 47

Half-cells, 535

**Half-life** The length of time it takes for one-half of a radioactive sample to decay, 567–571, 569r, 572r

carbon-14, 570

decay curve, 568

radioisotopes, examples, 569r

**Half-reaction method** A method of balancing oxidation–reduction reactions in which the half-reactions are balanced separately and then combined to give the complete reaction, 528–533

**Five elements** Examples of half-reactions balanced using, 528–533

Half-reactions, 522–523

**Halogen** An element in Group 7A (17), 138

Hand sanitizers, 610

Haworth structure, 644–646

Hazardous wastes, 578

HDL-cholesterol, 656

Heart attack, 662

**Heat** The energy associated with the motion of particles in a substance, 112

calculating in reaction, 297–298

calculations using specific heat, 115–116

of condensation, 333–334

specific, 115–118, 115r

units of, 112

**Heat equation** A relationship that calculates heat (q) given the mass, specific heat, and temperature change for a substance, 116–118

**Heating curve** A diagram that illustrates the temperature changes and changes of state of a substance as it is heated, 334–335

**Heat of fusion** The energy required to melt exactly 1 g of a substance at its melting point. For water, 334 J is needed to melt 1 g of ice; 334 J is released when 1 g of water freezes, 332–333

**Heat of reaction** The heat (symbol ΔH) absorbed or released when a reaction takes place at constant pressure, 295

**Heat of vaporization** The energy required to vaporize 1 g of a substance at its boiling point. For water, 2260 J is needed to vaporize exactly 1 g of water; 1 g of steam gives off 2260 J when it condenses, 333–334

Heliox, 103

Helix
double, 677–679

triple, 668, 669

Hemlock, 625

**Hemodialysis** A cleansing of the blood by an artificial kidney using the principle of dialysis, 427–428

Hemoglobin, 663r
carbon monoxide binding to, 268
carrying oxygen, 268, 663, 663r

function of, 663
Hemolysis: A swelling and bursting of red blood cells in a hypotonic solution due to an increase in fluid volume, 426

Henry's law: The solubility of a gas in a liquid is directly related to the pressure of that gas above the liquid, 400

Heroin: 625

Hess's law: Heat can be absorbed or released in a single chemical reaction or in several steps, 299

Heterogeneous equilibrium: An equilibrium system in which the components are in different states, 451

Heterogeneous mixture: 102

Hexane: 592r

Hexose: 643

Hibernation: 659

Hexose: 643

Histidine: 664

Hip replacement: 602

High-density polyethylene: 603

Hibernation: 659

Hexose: 643

Histidine: 664

Hornet: 81

Horizontal axis on line graph: 45

Hormone: 663r

Hot pack: 298

Human body: body fat, 76

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hair on scalp, 47

homeostasis and, 461

temperature scale on which the

Human body: body fat, 76

body temperature, 461

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homeostasis and, 461

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Human body: body fat, 76

body temperature, 461

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Human body: body fat, 76

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Human body: body fat, 76

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Human body: body fat, 76

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Human body: body fat, 76

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homeostasis and, 461

temperature scale on which the

Human body: body fat, 76

body temperature, 461

hair on scalp, 47

homeostasis and, 461

temperature scale on which the

Human body: body fat, 76

body temperature, 461

hair on scalp, 47

homeostasis and, 461

temperature scale on which the

Hydrocarbons: A type of organic compound that contains only carbon and hydrogen, 590

Hydrochloric acid (HCl): 497

Hydrogenation: 601–602, 656–657

Hydrogen bond: The attraction between a partially positive H atom and a strongly electronegative atom of N, O, or F, 329, 330r, 668 in DNA double helix, 678 in triple helix, 668, 669 in water, 392

Hydronium ion: $\text{H}_3\text{O}^+$: The ion formed by the attraction of a hydrogen ion, H+, to a water molecule, 477

Hydrophilic polar amino acid: 663

Hydrophobic nonpolar amino acid: 663

Hydroxide ion: 210

Hydroxyl group: 608–609, 616, 626 in alcohols, 608, 616 phenol, 608

Hyperbaric chambers: 412

Hyperglycemia: 644

Hyperthermia: 112

Hyperthyroidism: 572

Hypertonic solution: 426–427

Hypoglycemia: 644

Hypothermia: 111

Hypothesis: An unverified explanation of a natural phenomenon, 34

Hypotonic solution: 426

Hypoxia: 459

Ideal gas constant, R: A numerical value that relates the quantities P, V, n, and T in the ideal gas law, $PV = nRT$, 373

Ideal gas law: A law that combines the four measured properties of a gas: $PV = nRT$, 373–375

Imaging: CT scans, 574

PET and, 574

radioisotopes and, 562–563, 572r

Immunoglobulins: 663r

Inch: 59

Immunoglobulins: 663r

Inhalation: 361

Infrared radiation: 163

Induced-fit model: A model of enzyme action in which the shape of a substrate and the active site of the enzyme adjust to give an optimal fit, 673, 674

Inflammation: cold pack and, 298 reducing, 298

Infared radiation: 163

Inhalation: 361

Inorganic compound: 589, 590r

Insoluble ionic compound, 400–401, 400r, 401r

Inspiration: 361

Insulin: 663r, 668

International System of Units (SI): The official system of measurement throughout the world except for the United States that modifies the metric system, 58

International System of Units: The SI system used for organic compounds, 591

Iodine-125: 574

Isoeugenol: 611

Isoleucine: 664r

Isoleucine: 664r

Iron-59: 569

Irradiated foods: 565

Irradiated foods: 565

Isoeugenol: 611

Isocitric acid: 618

Isotonic solution: 426

Isotope: An atom that differs only in mass number from another atom of the same element. Isotopes have the same atomic number (number of protons) but different numbers of neutrons, 148–152

half-life, 567–571, 572r

radioactive, 554–557, 554r, 569r, 572–575, 572r

stable and radioactive, 554r

IUPAC: (International Union of Pure and Applied Chemistry) system: A naming system used for organic compounds, 591

alcohols, 609

aldehydes, 613

alkanes, 589–598, 592r

alkynes, 600

amines, 627

aromatic compounds, 606–607

carboxylic acids, 616–617 esters, 620 ethers, 610 ketones, 615

phenols, 609

J

Joule (J): The SI unit of heat energy; $4.184 \text{ J} = 1 \text{ cal}, 112, 295

K

Kelvin (K): temperature scale: A temperature scale on which the lowest possible temperature is 0 K, 60, 108

comparison of temperature scales, 107, 108r

Ketatin: 663r

α-Ketoglutaric acid: 618

Ketone: An organic compound in which a carbonyl group (C==O) is bonded to two carbon atoms, 600r, 612, 642–643

naming, 615
Ketose A monosaccharide that contains a ketone group, 642
Key math skills, 40–47, 492, 496
Kidney
dialysis, 427–428
glomerulus, 427
stones, 398–399
Kilo (k), 69, 70

Kilocalorie (kcal) An amount of heat energy equal to 1000 calories, 112
Kilogram (kg) A metric mass of 1000 g, equal to 2.205 lb. The kilogram is the SI standard unit of mass, 59
Kilojoule (kJ), 112, 295

Kinetic energy The energy of moving particles, 111

Kinetic molecular theory of gases A model used to explain the behavior of gases, 355–356

Krebs cycle, 618
See also Citric acid cycle

L

Labeling, food nutrition, 120
Laboratory technician, 473
Lactase A disaccharide consisting of glucose and galactose found in milk and milk products, 643, 648, 650

Lanolin, 658
Latin names for elements in clinical usage, 133–134
Lauric acid, 654, 655

Law of conservation of mass In a chemical reaction, the total mass of the reactants is equal to the total mass of the products; matter is neither lost nor gained, 280

LDL-cholesterol, 656
LD₃₀ (lethal dose), 82

Lead(II) oxide, 527

Lead storage battery, 538

Le Châtelier’s principle When a stress is placed on a system at equilibrium, the equilibrium shifts to relieve that stress, 456–462, 458r, 460

Length

equalities, 71, 74
measuring, 70–71
metric system, 59, 70–71
units of measurement, 59

Lethal dose (LD₃₀), 82

Leucine, 664r

Lewis structure A structure drawn in which the valence electrons of all the atoms are arranged to give octets except two electrons for hydrogen, 311
diatomic molecules, 311, 311r
double and triple bonds, 314–315
drawing, 311–314, 318r
exceptions to octet rule, 315
for hydrogen molecule, 311
for molecular compounds, 311, 312r
for molecules and polyatomic ions, 310–316

Lewis symbol The representation of an atom that shows valence electrons as dots around the symbol of the element, 182–183, 183r, 310, 310r

Light bulbs, energy-saving, 167–168
“Like dissolves like,” 393

Like terms, 43
Limestone, 377

Limiting reactant The reactant used up during a chemical reaction, which limits the amount of product that can form, 288–293

Calculating mass of product, 291–292

Calculating moles of product, 289–290

Percent yield, 293–295

Line-angle structural formula A formula which shows a zigzag line in which carbon atoms are represented as the ends of each line and as corners, 592–593

Linear The shape of a molecule that has two bonded atoms and no lone pairs, 319, 321

Linoleic acid, 655

Linolenic acid, 655

Lips A family of compounds that is nonpolar in nature and not soluble in water; includes fats, oils, waxes, and steroids, 654–662
types of, 654

See also Fat; Fatty acid

Lipprotein, 663

HDL and LDL, 656
See also Cholesterol

Liquid A state of matter that takes the shape of its container but has a definite volume, 104, 104r
density of common liquids, 85r
properties of, 104r

Solutions, 392

Liter (L) The metric unit for volume that is slightly larger than a quart, 59

Measuring volume, 71–72

Unit of measurement, 59

Lock-and-key model, 673

Lone pair, 311

Low-density polyethylene, 603

Low-pressure system, 358

Lysine, 664r, 665

M

Macrominerals, 140

Magnetic resonance imaging (MRI), 574

Malaria, 625

Malic acid, 619

Maltaise, 652

Maltose A disaccharide consisting of two glucose units; it is obtained from the hydrolysis of starch, 650

Margarine, 656

Mass A measure of the quantity of material in an object, 59–60

calculations for reactions, 285–288

Conservation of, 280–282

equalities, 72, 74r

Measuring, 72

Relation to moles and particles, 282–284

Units of measurement, 59–60

Mass number The total number of protons and neutrons in the nucleus of an atom, 146–148, 147r

Mass percent (m/m) The grams of solute in 100 g of solution, 235, 404–406

Mass percent composition The percent by mass of the elements in a formula, 235–236

Empirical formula calculated from, 239

Molar mass used to calculate, 236

Molar mass (M) A measure of the quantity of material in a sample, 59–60, 72

Measured numbers A number obtained when a quantity is determined by using a measuring device, 61–65

Measurement

body fat, 76

chemistry and, 58–98
conversion factors, 73–78

density, 84–90

Length, 59, 70–71

mass, 59–60, 72

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Pressure, 357–358, 358r

Problem solving, 78–84

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Temperature, 60

time, 60

Units of, 58–61, 58r

Volume, 59, 71–72

Volume displacement, 86–87

Weight, 59–60

Mega (M), 70r

Melamine, 243

Melting The change of state from a solid to a liquid, 332

Melting point (mp) The temperature at which a solid becomes a liquid (melts). It is the same temperature as the freezing point, 332

Attractive forces and, 329, 330r

Bond type and, 329r

Fat, 659, 660

Fatty acid, 655r

Inorganic compound, 590r

Oil, 659, 660

Organic compound, 590r

Mendeleev, Dmitri, 136

Messenger RNA. See mRNA

Metabolic acidosis/alkalosis, 505–506, 506

Metal An element that is shiny, malleable, ductile, and a good conductor of heat and electricity. The metals are located to the left of the heavy zigzag line on the periodic table, 138–140

Acids and, 498

Activity series for, 534r

Alkaline earth metals, 138

Alkaline earth metals, 138

Characteristics of, 139r

Forming positive ion, 205–206, 206r, 207r

Ionic charge, 205–207

Ions, common, 595r

Oxidation of, 539–540

In periodic table, 137

Variable charge, 205–207

Metallic character A measure of how easily an element loses a valence electron, 185–186, 186r

Metalloid Elements with properties of both metals and nonmetals located along the heavy zigzag line on the periodic table, 138–140, 139r
Methanamide, 627

The metric unit for length that is
Metastable, 562

The mass in grams of 1 mol of an
Molar mass

The number of moles of solute
Molarity (M)

The physical combination of two or
Mixture

Mineral oil, 597

Millimeter (mm), 70

Milli

Milk sugar (lactose), 648

Microwaves, 163

Microscope, 143

Microminerals, 140

μ

Micro

Nitracine, 625

Nitrile, 210r

Nitrogen-containing base, 674

Nitrogen dioxide, 279, 530–531

Nitrogen narcosis, 103

Nitrox, 103

Noble gas An element in Group 8A (18) of the
periodic table, 138, 196

Nonelectrolyte A substance that dissolves in
water as molecules; its solution does not
conduct an electrical current, 395, 395r

Nonmetal An element with little or no luster
that is a poor conductor of heat and
electricity. The nonmetals are located to
the right of the heavy zigzag line on the
periodic table, 138–139, 139r

bonding pattern, 311, 311r

polyatomic ion, 209–213

Nonpolar amino acids, 663, 664r

Nonpolar covalent bond A covalent bond
in which the electrons are shared
equally between atoms, 324

Nonpolar molecule A molecule that has only
nonpolar bonds or in which the bond
dipoles cancel, 327

Nonpolar solutes, 394

N-terminal amino acid, 666

Nuclear chemistry, 553–585

half-life of radioisotopes, 567–571

medical applications, 572–575

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nuclear fission and fusion, 575–578

nuclear reactions, 558–564

radiation measurement, 564–567

See also Radiation

Nuclear equation, 558

Nuclear fission, 575–578, 577

Nuclear fusion, 576–578

Nuclear medicine, 572–575

Nuclear power plant, 578–579

Chernobyl, 566

Nuclear reactions, 558–564

Nucleic acid A large molecule composed of
nucleotides; found as a double helix in DNA
and as the single strands of RNA, 674–680
bases in, 674–675
components of, 674–680
nucleotide structure, 674–680
structure of, 676–677

Nucleoside The combination of a pentose
sugar and a base, 675

producing, 675

Nucleotide Building block of a nucleic acid
consisting of a base, a pentose sugar
(ribose or deoxyribose), and a phosphate
group, 675–676

in DNA, 675–676

forming, 675–676

name in DNA and RNA, 676

naming, 676

in RNA, 675–676

structure, 674, 676

Nucleus The compact, extremely dense center
of an atom, containing the protons and
neutrons of the atom, 143

changing due to radiation, 560

fusion, 575

splitting, 575

Numerator, 73

NutraSweet, 650
Nutrition, 119–122
   caloric value of foods, 119–121, 120r
Nutrition Facts label, 120

**Observation** Information determined by noting and recording a natural phenomenon, 34

**Octet** A set of eight valence electrons, 196

**Octet rule** Elements in Groups 1A to 7A (1, 2, 13 to 17) react with other elements by forming ionic or covalent bonds to produce a stable electron configuration, usually eight electrons in the outer shell, 196 exceptions, 315

Octyl ethanoate, 621r

**Oil** A triacylglycerol that is usually a liquid at room temperature and is obtained from a plant source, 658
   hydrogenation, 656–657
   melting point, 659–661
   polyunsaturated, 659
   at room temperature, 660
   vegetable, 656
   See also Crude oil

Oil spills, 598

Oleic acid, 602, 655r, 656

Omega-3 fatty acids in fish oils, 685

Opium, 625

**Orbital** The region around the nucleus where electrons of certain energy are most likely to be found: s orbitals are spherical; p orbitals have two lobes, 169
   capacity, 171
   shape of, 169–171

**Orbital diagram** A diagram that shows the distribution of electrons in the orbitals of the energy levels, 172–174
   Period 1, 173–174
   Period 2, 174–175
   Period 3, 175–176

Organic compounds, 589, 590
   bonding, 590r
   carbon compounds, 590–591
   classification of, 599, 600r
   functional groups, 599–605
   naming, 591, 592r, 593, 595–596
   properties of, 590r

Osmosis The flow of a solvent, usually water, through a semipermeable membrane into a solution of higher solute concentration, 425–426
   properties of solutions, 425–429

**Osmotic pressure** The pressure that prevents the flow of water into the more concentrated solution, 425–426

Osteoporosis, 87

Overweight, 99

Oxaloacetic acid, 618–619

Oxidation The loss of electrons by a substance. Oxidation may involve the addition of oxygen or the loss of hydrogen, 269–272, 522

**Oxidation number** A number equal to zero in an element or the charge of a monatomic ion; in molecular compounds and polyatomic ions, oxidation numbers are assigned using a set of rules, 523, 524r
   assigning of, 524r
   oxidation-reduction identified using, 525–526

**Oxidation-reduction reaction** A reaction in which electrons are transferred from one reactant to another, 269–272
   balancing of, 528–533
   in biological systems, 270–271
   characteristics of, 271, 271r
   definition of, 269
   electrical energy and, 533–540
   half-reactions, 522–523, 528–530
   oxidation numbers used to identify, 525–526
   writing, 522

   zinc and copper(II) sulfate, 270, 271

**Oxidizing agent** The reactant that gains electrons and is reduced, 526

Oxycodone (OxyContin), 625

Oxygen-15, 569

Oxyn–hemoglobin equilibrium, 459

Ozone, 316

P Pain relievers, 625

Palmitic acid, 655, 657

Palladium-103, 572

Pain relievers, 625

Palmitoleic acid, 655

Palmitoleic acid, 655

Palladium-103, 572

Pain relievers, 625

Pentyl ethanoate, 621

Pentyl butanoate, 621

Pentose sugars, 675

Phenylalanine, 650, 664

Phenolphthalein, 500

Polar acid An acid that is soluble in water, 136–141, 137

Periodic table
   Periodic properties, 181–186
   Period
   Percent yield
   Percent concentration, 404–406, 410r

**Percent yield** The ratio of the actual yield for a reaction to the theoretical yield possible for the reaction, 293–295

Period A horizontal row of elements in the periodic table, 136

Periodic properties, 181–186

**Periodic table** An arrangement of elements by increasing atomic number such that elements having similar chemical behavior are grouped in vertical columns, 136–141, 137
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*Poly* as a suffix, 599

Polyatomic ion
   A group of covalently bonded nonmetal atoms that has an overall electrical charge, 209–213
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   molecular shapes and, 318–323
   name of, 209–210, 210r, 212
   products containing, 209
   writing formula, 210–211

Polychloroethylene (Saran), 603, 604r

Polyethylene, 603, 603, 604r

Polymer A large molecule that is composed of many small, repeating structural units that are identical, 599, 602–605
   alkenes, 599, 602–605, 604r
   definition of, 602
   plastics, recycle, 603
   recycling, 603
   synthetic, 602–603
   Polypropylene, 603, 604r

Poly saccharide A polymer of many monosaccharide units, usually glucose. Polysaccharides differ in the types of glycosidic bonds and the amount of branching in the polymer, 647–654
   cellulose, 652–653
   definition of, 651–653
   digesting, 652

Indicators, 492–495, 504–507

Indicators, 500

scale, 491

Phenol, 606, 610–611

Phenylalanine, 650, 664r

Phenylketonuria (PKU), 650

Phosphodiester linkage
   The phosphate link that joins the hydroxyl group in one nucleotide to the phosphate group on the next nucleotide, 676

Phosphorus-32, 572r

Photon A packet of energy that has both particle and wave characteristics and travels at the speed of light, 165

Photosynthesis, 642

Physical change A change in which the physical properties of a substance change but its identity stays the same, 104–105, 105r

**Physical properties** The properties that can be observed or measured without affecting the identity of a substance, 104, 104r, 105r

Pico (p), 70r

PKU (phenylketonuria), 650

Place values, 40–46

Plaque, 662

Plasma, 393

Plastics, 603–604

Recycling, 603

pOH A measure of the [OH−] in a solution; pOH = log[OH−], 495

Polar amino acid
   An amino acid that is charged, 664

Polar covalent bond
   A bond in which electrons are shared unequally, 603

Polar covalent bond
   A bond in which the electrons are shared unequally between atoms, 324

Polarity A measure of the unequal sharing of electrons, indicated by the difference in electronegativities, 324–326

Polar molecules, 327–328

Polar molecule A molecule containing bond dipoles that do not cancel, 327–328

Polar solute, 394

Polar solvent, 392

Polyatomic ion
   A group of covalently bonded nonmetal atoms that has an overall electrical charge, 209–213
   compounds containing, 210, 212r
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   name of, 209–210, 210r, 212
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Polydichloroethylene (Saran), 603, 604r

Polyethylene, 603, 603, 604r

Polymer A large molecule that is composed of many small, repeating structural units that are identical, 599, 602–605
   alkenes, 599, 602–605, 604r
   definition of, 602
   plastics, recycle, 603
   recycling, 603
   synthetic, 602–603
   Polypropylene, 603, 604r

Polysaccharide A polymer of many monosaccharide units, usually glucose. Polysaccharides differ in the types of glycosidic bonds and the amount of branching in the polymer, 647–654
   cellulose, 652–653
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**Positron** A particle of radiation with no mass and a positive charge, symbol $e^+$ or $\phi^+$, produced when a proton is transformed into a neutron and a positron, 555, 555

Potassium-40, 569

**Potential energy** A type of energy related to position or composition of a substance, 111–112

Pound, conversion factor, 74

Principle quantum number ($n$) The number that specifies the size of the measurement. All prefixes are related on a decimal scale, 69–73, 70r, 623

**Pressure** The force exerted by gas particles that hit the walls of a container, 355

Principal quantum number ($n$) The number ($n = 1, n = 2 \ldots$) assigned to an energy level, 166

PRISM, light passing through, 162

Problem solving, 78–84

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Products The substances formed as a result of a chemical reaction, 255–256

Proline, 664r

Propane, 267, 589, 590, 592r

Propanediol (glycerol, glycerin), 610–611

Propanoic acid, 616, 617r

Propionic acid, 616

Propionylc acid, 616

Propyl ethanoate, 621r

Prostate cancer, 574–575

**Protein** A term used for biologically active polypeptides that have many amino acids linked together by peptide bonds, 663–668
classification of, 663r
composition of, 663

definition of, 668
energy content, 120r
as enzymes, 672–674
primary structure, 668
secondary structure, 668, 669
structure of, 668–672, 671r
20 amino acids in, 665

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translation, 682, 683

Proton A positively charged subatomic particle having a mass of about 1 amu and found in the nucleus of an atom; its symbol is $p$ or $p^+$, 143, 144, 145r, 555r

psi (pounds per square inch), 355, 357, 358

Pyruvic acid, 618

Rad (radiation absorbed dose) A measure of an amount of radiation absorbed by the body, 565, 566r

Radiation Energy or particles released by radioactive atoms, 554

average annual per person, 567r

background, 566, 566r

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medical applications, 572–575

protecting from, 556–557

sickness, 567

therapy, 573–574

treating food, 565
types of, 555

Radioactive, 554

Radioactive decay The process by which an unstable nucleus breaks down with the release of high-energy radiation, 558–564

Radioactive iodine, 572

Radioactive iodine uptake, 573

Radioactive isotope, 554–557, 554r

producing, 562–564

Radioactive waste, 578

Radioactivity, 572–575

Radioisotope A radioactive atom of an element, 554–557, 554r

half-life, 567–571, 572r

medical applications, 569r, 572r, 573–574

scans with, 573–574

Radiologist, 553

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Radium-226, 557r, 559, 569r

Radon-222, 566, 569r

Rate of Reaction The speed at which reactants form products, 440–445

factors affecting, 442–443, 443r

formulas, 442

Reactants The initial substances that undergo change in a chemical reaction, 255–256

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Reaction catalysts and, 443, 443r, 673
direction of, 483–484

heat of, 295

See also Chemical reaction; Rate of reaction

Redox. See Oxidation–reduction reaction

Reducing agent The reactant that loses electrons and is oxidized, 526–527

Reduction The gain of electrons by a substance. Reduction may involve the loss of oxygen or the gain of hydrogen, 269–272, 522

Registered nurse, 57

Rem (radiation equivalent in humans) A measure of the biological damage caused by the various kinds of radiation (rad $\times$ radiation biological factor), 565, 566r

Replication The process of duplicating DNA by pairing the bases on each parent strand with their complementary bases, 679–680

Representative element An element in the first two columns on the left of the periodic table and the last six columns on the right that has a group number of 1A through 8A or 1, 2, and 13 through 18, 136–137

Repulsion and attraction, 143

Resonance structures Two or more Lewis structures that can be drawn for a molecule or polyatomic ion by placing a multiple bond between different atoms, 316–318

Respiration, 642

Respiratory acidosis/alkalosis, 505–506, 506r

Respiratory therapist, 353

Reverse osmosis, 426

Reverse reaction, 445, 446

Reversible reaction A reaction in which a forward reaction occurs from reactants to products, and a reverse reaction occurs from products back to reactants, 445, 446

Ribonucleic acid, See RNA

Ribose, 242r, 675

Ribosomal RNA (rRNA), 680, 680r

Ribosome, 682

Risk–benefit assessment, 82

RNA Ribonucleic acid; a type of nucleic acid that is a single strand of nucleotides containing ribose, phosphate, and the four bases: adenosine, cytosine, guanine, and uracil, 674–684

bases in, 674–675

geneic code and, 681–682, 682r

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S

Semipermeable membrane, 419
Serine, 664
Shapes of molecules, 318–323, 321r
Shielding Materials used to provide protection from radioactive sources, 556–557
SI See International System of Units (SI)
Sievert (Sv) A unit of biological damage (equivalent dose) equal to 100 rem, 565, 566r
Significant figures (SFs) The numbers recorded in a measurement, 62–69 in calculation, 65–69 in measured numbers, 62t multiplication and division, 66 rounding off, 65–66 significant zeros, 62–63, 62r, 67–68
Single replacement reaction A reaction in which an element replaces a different element in a compound, 265–266, 267t
Solid A state of matter that has its own shape and volume, 103, 104t density calculations, 85t, 86–87 formation of, 401–402 properties of, 104t as solution, 392t
Solubility The maximum amount of solute that can dissolve in 100. g of solvent, usually the component present in greater amount, 391–394, 391t
Solving algebraic equations, 43–44
Specific gravity (sp gr) A relationship between the density of a substance and the density of water, 88–89
Specific heat (SH) A quantity of heat that changes the temperature of exactly 1 g of a substance by exactly 1 °C, 115–118, 115t
Standard number, 50
Standard temperature and pressure (STP), 371–372
Starch, 651
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Start codon, 682, 682r
States of matter Three forms of matter: solid, liquid, and gas, 103–106, 104t
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nuclear power plants and, 578–579
Stearic acid, 602, 655r
Steroids Types of lipid composed of a multicyclic ring system, 662
Stimulant caffeine, 82, 625
central nervous system, 625
nicotine, 625
Stomach acid, 497
STP Standard conditions of exactly 0 °C (273 K) temperature and 1 atm pressure used for the comparison of gases, 371–372
Stress on equilibrium, 456–458, 458t

**Strong acid** An acid that completely dissociates in water, 480–483, 480t, 486t

**Strong base** A base that completely dissociates in water, 482–483

**Strong electrolyte** A compound that ionizes completely when it dissolves in water; its solution is a good conductor of electricity, 395, 395t

Strontium-90, 569

Strontium-85, 572

A base that completely dissociates in water, 480–483, 480t

Strong acid

Strong electrolyte

**Subatomic particle** A particle within an atom; protons, neutrons, and electrons are subatomic particles, 143, 145t

Sublevel A group of orbitals of equal energy within energy levels. The number of sublevels in each energy level is the same as the principal quantum number (n), 168–169, 171t

block order, exceptions in, 180

electron configurations using, 172–176

**Sublimation** The change of state in which a solid is transformed directly to a gas without forming a liquid first, 338

Subscripts in formulas, 203

**Substituent** A group of atoms such as an alkyl group or a halogen bonded to the alkyl group itself, 168–169, 171t

equalities, 74

equations, 74

**Subtraction** The maximum amount of product that a reaction can produce from a given amount of reactant, 293

**T** Technetium (Tc), 562, 562

Technetium-99m, 562, 569t, 572t

Teeth, 521

Teflon, 603, 604t

**Temperature** An indicator of the hotness or coldness of an object, 60, 107–111

body, 110–111, 461

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reaction rate and, 442, 443t

solubility and, 399

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units of measurement, 60

volume and (Charles’s law), 362–364

Temperature scale. See Celsius (°C) temperature scale; Fahrenheit (°F) temperature scale; Kelvin (K) temperature scale

Tetrahydrofuran (THF), 604t

**Tetrahedral** The shape of a molecule with four bonded atoms, 320, 321t

**Theoretical yield** The maximum amount of product that a reaction can produce from a given amount of reactant, 293

**Theory** An explanation for an observation supported by additional experiments that confirm the hypothesis, 34–35

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Brunsted-Lowry, 477–479

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Thymine (T), 674–675

Thymol, 611

Thyroid gland, hyperactive, 572–573

Thyroid scan, 573

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equalities, 74t

unit of measurement (second), 60

**Titration** The addition of base to an acid sample to determine the concentration of the acid, 500–502

**Trace elements in body** 140, 141

**Trans fatty acid** 656–657

**Transfer RNA** See tRNA

**Transition element** An element in the center of the periodic table that is designated with the letter “B” or the group number of 3 through 12, 137

**Translation** The interpretation of the codons in mRNA as amino acids in a peptide, 681, 682, 683

mRNA and, 680, 683

protein, 680, 683

**Transmutation**, 560

**Triacylglycerols** A family of lipids composed of three fatty acids bonded through ester bonds to glycerol, a trihydroxy alcohol, 658–662

hydrogenation, 656–657

reactions of, 661–662

See also Fat

**Triglyceride**, 658. See also Triacylglycerol

**Trigonal planar** The shape of a molecule with three bonded atoms and no lone pairs, 319, 321t

**Trigonal pyramidal** The shape of a molecule that has three bonded atoms and one lone pair, 321t, 342

**Trimethylamine**, 623

**Trimix**, 103

**Trinitrotoluene (TNT)**, 607

**Triple bond** A sharing of three pairs of electrons by two atoms, 314

**Tritelecular acid** 661–662

**Tryptase**, 664t

**Tryptophan**, 664t

U

**Ultraviolet light**, 163

**Unit of measurement**, 58–61, 58t

See also SI unit

**Unsaturated fat**

hydrogenation of, 296, 656–657

oxidation, 661–662

vegetable oil, 656

**Unsaturated fatty acid**, 656

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Unsaturated hydrocarbon, 599

**Unsaturated solution** A solution that contains less solute than can be dissolved, 397

Uracil (U), 674–675

Uranium-235, 575

Uranium-238, 558, 568

Uric acid, 398

Uridine monophosphate (UMP), 676

Urine

pH of, 491

water loss from, 393

U.S. System, 74–75

V

Valence, variable, 205–207

**Valence electrons** The electrons in the highest
energy level of an atom, 181–182, 182t, 186r
bonding and, 196
definition of, 181
sharing, 213–217

**Valence shell electron-pair repulsion (VSEPR) theory** A theory that predicts the shape of a molecule by moving the electron groups on a central atom as far apart as possible to minimize the mutual repulsion of the electrons, 318–323

Valine, 664t
Vanilla, 614
Vanillin, 607, 611
Vaporization, 333–334

**Vapor pressure** The pressure exerted by the particles of vapor above a liquid, 366–367, 367t

Variable charge, 205–207

Vegetable oils, 602, 656, 659–660
Vertical axis, on line graph, 45
Veterinarian, 224
Vinegar, 616
Visible light, 163

**Voltaic cell** A type of cell with two compartments that uses spontaneous oxidation–reduction reactions to produce electrical energy, 535–537

**Volume (V)** The amount of space occupied by a substance, 59
change, equilibrium and, 458–459, 460t
description of, 59

displacement, in calculating solid density, 86–87
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**Volume percent (v/v)** The milliliters of solute in 100 mL of solution, 406–407

**W**

Water boiling and freezing points, 107–111, 108t, 366, 366r
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in human body, 393
as solvent, 392
specific heat, 115–116, 115t
vapor pressure, 366r, 366–367

**Water dissociation expression, \(K_w\)** The product of [H\(_3\)O\(^+\)] and [OH\(^-\)] in solution;
\[ K_w = [H_3O^+][OH^-] \]

Watson, James, 677

**Wavelength** The distance between adjacent crests or troughs in a wave, 162–163

Wax, 658t, 658

**Weak acid** An acid that is poor donor of H\(^+\) and dissociates only slightly in water, 480–481, 480r, 485t
dissociation constants for, 484–487, 486t

**Weak base** A base that is poor acceptor of H\(^+\) and produces only a small number of ions in water, 482–487

**Weak electrolyte** A substance that produces only a few ions along with many molecules when it dissolves in water; its solution is a weak conductor of electricity, 395, 395t

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<thead>
<tr>
<th>Length</th>
<th>SI Unit Meter (m)</th>
<th>Volume</th>
<th>SI Unit Cubic Meter (m³)</th>
<th>Mass</th>
<th>SI Unit Kilogram (kg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 meter (m) = 100 centimeters (cm)</td>
<td>1 liter (L) = 1000 milliliters (mL)</td>
<td>1 kilogram (kg) = 1000 grams (g)</td>
<td>1 mg = 1000 mcg</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 m = 1000 millimeters (mm)</td>
<td>1 mL = 1 cm³</td>
<td>1 g = 1000 milligrams (mg)</td>
<td>1 lb = 16 oz</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 cm = 10 mm</td>
<td>1 L = 1.057 quart (qt)</td>
<td>1 kg = 2.205 lb</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 kilometer (km) = 0.6214 mile (mi)</td>
<td>1 qt = 946.4 mL</td>
<td>1 lb = 453.6 g</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 inch (in.) = 2.54 cm (exact)</td>
<td>1 gallon (gal) = 3.785 L</td>
<td>1 mol = 6.022 × 10²³ particles</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

### Temperature | SI Unit Kelvin (K) | Pressure | SI Unit Pascal (Pa) | Energy | SI Unit Joule (J) |
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>T_F = 1.8(T_C) + 32</td>
<td>1 atm = 760 mmHg</td>
<td>1 atm = 101.325 kPa</td>
<td>1 calorie (cal) = 4.184 J (exact)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>T_C = (T_F - 32) / 1.8</td>
<td>1 atm = 760 mmHg</td>
<td>1 kcal = 1000 cal</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>T_K = T_C + 273</td>
<td>1 atm = 760 Torr</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

### Metric and SI Prefixes

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Power of Ten</th>
</tr>
</thead>
<tbody>
<tr>
<td>Prefixes That Increase the Size of the Unit</td>
<td></td>
<td></td>
</tr>
<tr>
<td>peta</td>
<td>P</td>
<td>10¹⁵</td>
</tr>
<tr>
<td>tera</td>
<td>T</td>
<td>10¹²</td>
</tr>
<tr>
<td>giga</td>
<td>G</td>
<td>10⁹</td>
</tr>
<tr>
<td>mega</td>
<td>M</td>
<td>10⁶</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>10³</td>
</tr>
<tr>
<td>Prefixes That Decrease the Size of the Unit</td>
<td></td>
<td></td>
</tr>
<tr>
<td>deci</td>
<td>d</td>
<td>10⁻¹</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>10⁻²</td>
</tr>
<tr>
<td>milli</td>
<td>m</td>
<td>10⁻³</td>
</tr>
<tr>
<td>micro</td>
<td>μ (mc)</td>
<td>10⁻⁶</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>10⁻⁹</td>
</tr>
<tr>
<td>pico</td>
<td>p</td>
<td>10⁻¹²</td>
</tr>
<tr>
<td>femto</td>
<td>f</td>
<td>10⁻¹⁵</td>
</tr>
</tbody>
</table>

### Formulas and Molar Masses of Some Typical Compounds

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Molar Mass (g/mol)</th>
<th>Name</th>
<th>Formula</th>
<th>Molar Mass (g/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonia</td>
<td>NH₃</td>
<td>17.03</td>
<td>Hydrogen chloride</td>
<td>HCl</td>
<td>36.46</td>
</tr>
<tr>
<td>Ammonium chloride</td>
<td>NH₄Cl</td>
<td>53.49</td>
<td>Iron(III) oxide</td>
<td>Fe₂O₃</td>
<td>159.70</td>
</tr>
<tr>
<td>Ammonium sulfate</td>
<td>(NH₄)₂SO₄</td>
<td>132.15</td>
<td>Magnesium oxide</td>
<td>MgO</td>
<td>40.31</td>
</tr>
<tr>
<td>Bromine</td>
<td>Br₂</td>
<td>159.80</td>
<td>Methane</td>
<td>CH₄</td>
<td>16.04</td>
</tr>
<tr>
<td>Butane</td>
<td>C₂H₁₀</td>
<td>58.12</td>
<td>Nitrogen</td>
<td>N₂</td>
<td>28.02</td>
</tr>
<tr>
<td>Calcium carbonate</td>
<td>CaCO₃</td>
<td>100.09</td>
<td>Oxygen</td>
<td>O₂</td>
<td>32.00</td>
</tr>
<tr>
<td>Calcium chloride</td>
<td>CaCl₂</td>
<td>110.98</td>
<td>Potassium carbonate</td>
<td>K₂CO₃</td>
<td>138.21</td>
</tr>
<tr>
<td>Calcium hydroxide</td>
<td>Ca(OH)₂</td>
<td>74.10</td>
<td>Potassium nitrate</td>
<td>KNO₃</td>
<td>101.11</td>
</tr>
<tr>
<td>Calcium oxide</td>
<td>CaO</td>
<td>56.08</td>
<td>Propane</td>
<td>C₃H₈</td>
<td>44.09</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td>CO₂</td>
<td>44.01</td>
<td>Sodium chloride</td>
<td>NaCl</td>
<td>58.44</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl₂</td>
<td>70.90</td>
<td>Sodium hydroxide</td>
<td>NaOH</td>
<td>40.00</td>
</tr>
<tr>
<td>Copper(II) sulfide</td>
<td>CuS</td>
<td>95.62</td>
<td>Sulfur trioxide</td>
<td>SO₃</td>
<td>80.07</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H₂</td>
<td>2.016</td>
<td>Water</td>
<td>H₂O</td>
<td>18.02</td>
</tr>
</tbody>
</table>