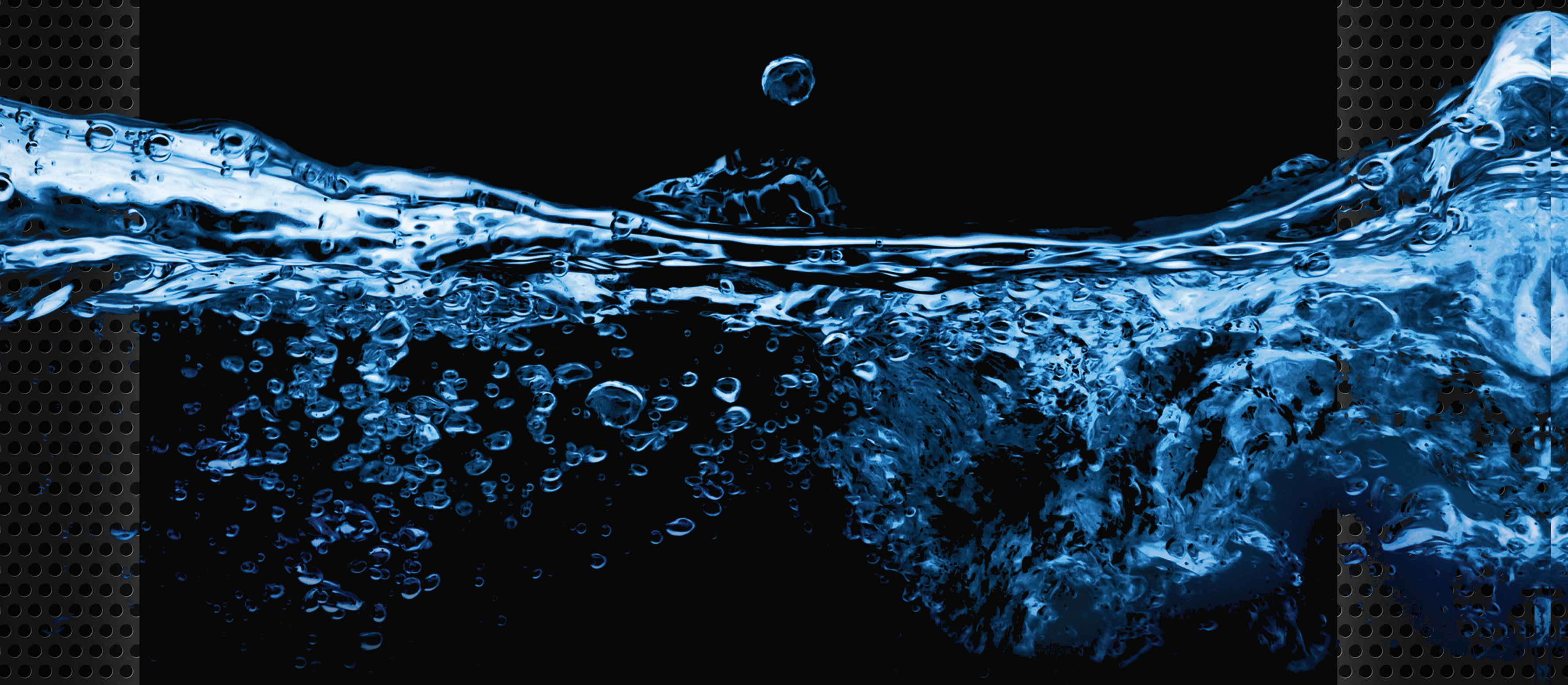
Chemistry

Fourth Edition



Julia Burdge

McGraw Hill Education

Chemistry

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COLLEGE OF WESTERN IDAHO

McGraw Hill Education

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About the Author

Julia Burdge received her Ph.D. (1994) from the University of Idaho in Moscow, Idaho. Her research and dissertation focused on instrument development for analysis of trace sulfur compounds in air and the statistical evaluation of data near the detection limit.

In 1994, she accepted a position at The University of Akron in Akron, Ohio, as an assistant professor and director of the Introductory Chemistry program. In the year 2000, she was tenured and promoted to associate professor at The University of Akron on the merits of her teaching, service, and research in chemistry education. In addition to directing the general chemistry program and supervising the teaching activities of graduate students, she helped establish a future-faculty development program and served as a mentor for graduate students and post-doctoral associates. Julia has recently relocated back to the northwest to be near family. She lives in Boise, Idaho, and holds an adjunct faculty position at the College of Western Idaho in Nampa.

In her free time, Julia enjoys precious time with her three children, and with Erik Nelson, her partner and best friend.

To the people who will always matter the most: Katie, Beau, and Sam.

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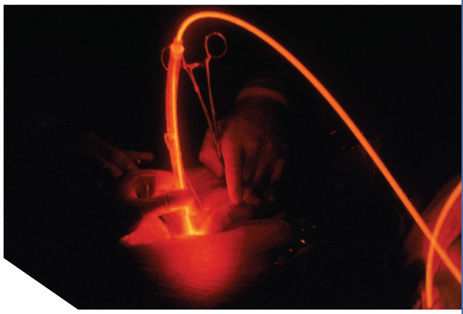

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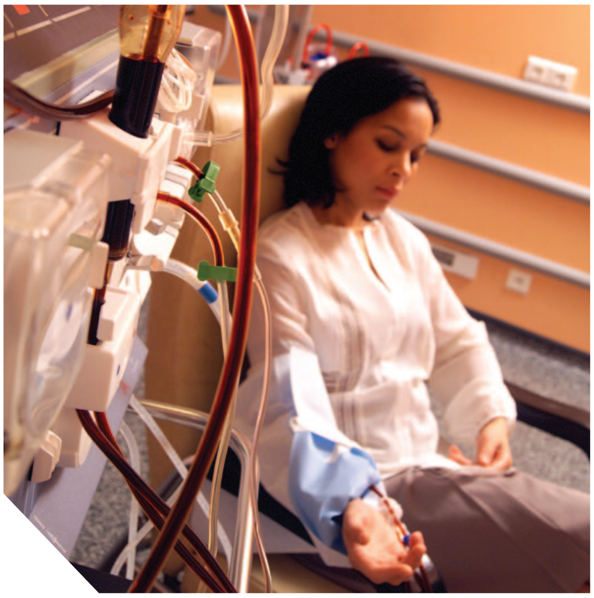

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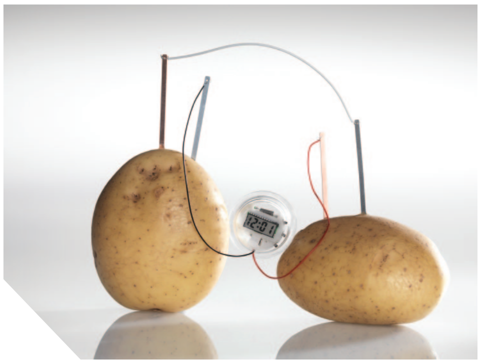

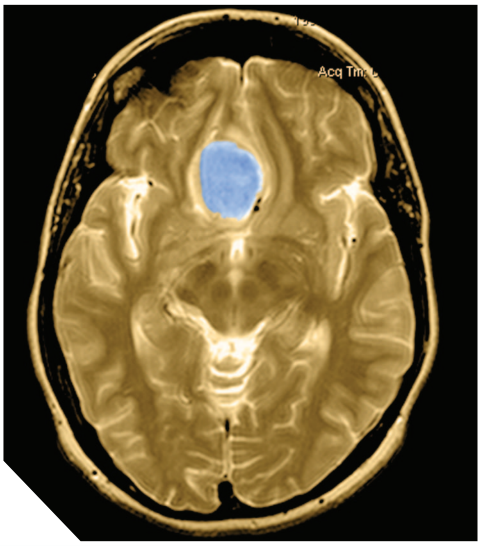

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Preface

Welcome to the exciting and dynamic world of Chemistry! My desire to create a general chemistry textbook grew out of my concern for the interests of students and faculty alike. Having taught general chemistry for many years, and having helped new teachers and future faculty develop the skills necessary to teach general chemistry, I believe I have developed a distinct perspective on the common problems and misunderstandings that students encounter while learning the fundamental concepts of chemistry—and that professors encounter while teaching them. I believe that it is possible for a textbook to address many of these issues while conveying the wonder and possibilities that chemistry offers. With this in mind, I have tried to write a text that balances the necessary fundamental concepts with engaging real-life examples and applications, while utilizing a consistent, step-by-step problem-solving approach and an innovative art and media program.

Key Features

Problem Solving Methodology

Sample Problems are worked examples that guide the student step-by-step through the process of solving problems. Each Sample Problem follows the same four-step method: Strategy, Setup, Solution, and Think About It (check).

Strategy: plan is laid out for solving the problem.

Setup: necessary information is gathered and organized.

Solution: problem is worked out.

Think About It:
– Assess the result.
– Provides information that shows the relevance of the result or the technique.
– Sometimes shows an alternate route to the same answer.

SAMPLE PROBLEM 4.8

For an aqueous solution of glucose ($C_6H_{12}O_6$), determine (a) the molarity of 2.00 L of a solution that contains 50.0 g of glucose, (b) the volume of this solution that would contain 0.250 mol of glucose, and (c) the number of moles of glucose in 0.500 L of this solution.

Strategy Convert the mass of glucose given to moles, and use the equations for interconversions of *M*, liters, and moles to calculate the answers.

Setup The molar mass of glucose is 180.2 g.

$$\text{moles of glucose} = \frac{50.0\text{ g}}{180.2\text{ g/mol}} = 0.277\text{ mol}$$

Solution (a) $\text{molarity} = \dfrac{0.227\text{ mol } C_6H_{12}O_6}{2.00\text{ L solution}} = 0.139\ M$

A common way to state the concentration of this solution is to say, "This solution is 0.139 *M* in glucose."

(b) $\text{volume} = \dfrac{0.250\text{ mol } C_6H_{12}O_6}{0.139\ M} = 1.80\text{ L}$

(c) moles of $C_6H_{12}O_6$ in 0.500 L $= 0.500\text{ L} \times 0.139\ M = 0.0695$ mol

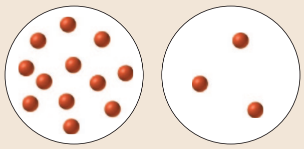
THINK ABOUT IT

Check to see that the magnitudes of your answers are logical. For example, the mass given in the problem corresponds to 0.277 mol of solute. If you are asked, as in part (b), for the volume that contains a number of moles smaller than 0.277, make sure your answer is smaller than the original volume.

Practice Problem ATTEMPT For an aqueous solution of sucrose ($C_{12}H_{22}O_{11}$), determine (a) the molarity of 5.00 L of a solution that contains 235 g of sucrose, (b) the volume of this solution that would contain 1.26 mol of sucrose, and (c) the number of moles of sucrose in 1.89 L of this solution.

Practice Problem BUILD For an aqueous solution of sodium chloride (NaCl), determine (a) the molarity of 3.75 L of a solution that contains 155 g of sodium chloride, (b) the volume of this solution that would contain 4.58 mol of sodium chloride, and (c) the number of moles of sodium chloride in 22.75 L of this solution.

Practice Problem CONCEPTUALIZE The diagrams represent solutions of two different concentrations. What volume of solution 2 contains the same amount of solute as 5.00 mL of solution 1? What volume of solution 1 contains the same amount of solute as 30.0 mL of solution 2?



solution 1　　solution 2

Each Sample Problem is followed by my ABC approach of three Practice Problems: **A**ttempt, **B**uild, and **C**onceptualize.

Practice Problem **A** (or "**A**ttempt") asks the student to apply the same Strategy to solve a problem very similar to the Sample Problem. In general, the same Setup and series of steps in the Solution can be used to solve Practice Problem A.

ATTEMPT

Practice Problem **B** (or "**B**uild") assesses mastery of the same skills as those required for the Sample Problem and Practice Problem A, but everywhere possible; Practice Problem B cannot be solved using the same Strategy used for the Sample Problem and for Practice Problem A. This provides the student an opportunity to develop a strategy independently, and combats the tendency that some students have to want to apply a "template" approach to solving chemistry problems. Practice Problems "Attempt" and "Build" have been incorporated into the problems available in Connect (R) and can be used in online homework and/or quizzing.

BUILD

Practice Problem **C** (or "**C**onceptualize") provides an exercise that probes the student's conceptual understanding of the material. Practice Problems C often include concept and molecular art.

CONCEPTUALIZE

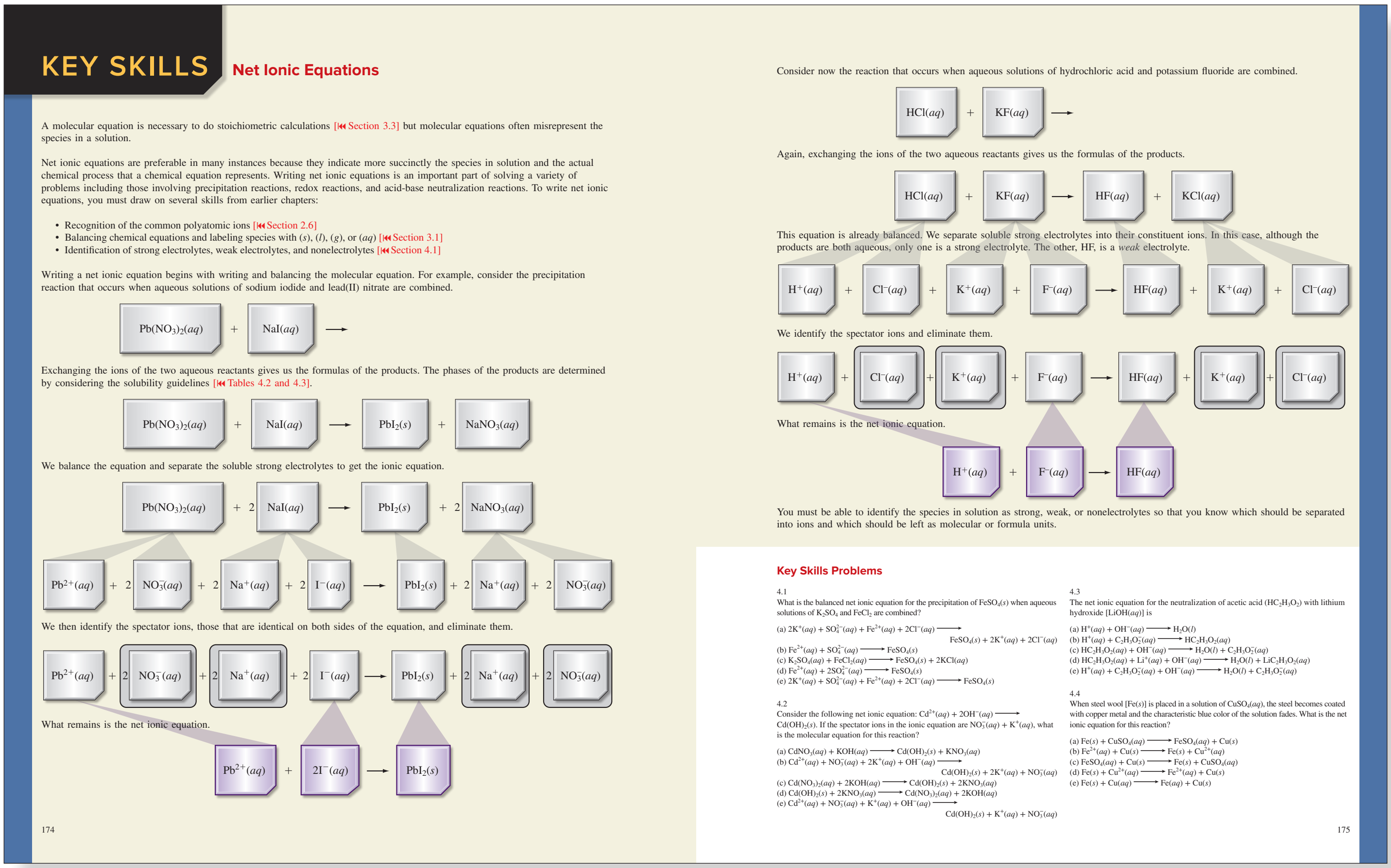
Each chapter's end-of-chapter questions and problems begin with an **Integrative Problem,** entitled *Applying What You've Learned.* These integrative problems incorporate multiple concepts from the chapter, with each step of the problem providing a specific reference to the appropriate Sample Problem in case the student needs direction.

> **Applying What You've Learned**
>
> Sports drinks typically contain sucrose ($C_{12}H_{22}O_{11}$), fructose ($C_6H_{12}O_6$), sodium citrate ($Na_3C_6H_5O_7$), potassium citrate ($K_3C_6H_5O_7$), and ascorbic acid ($H_2C_6H_6O_6$), among other ingredients. (a) Classify each of these ingredients as a nonelectrolyte, a weak electrolyte, or a strong electrolyte [⏮ Sample Problem 4.1]. (b) If a sports drink is 0.0015 *M* in both potassium citrate and potassium phosphate, what is the overall concentration of potassium in the drink [⏮ Sample Problem 4.11]? (c) The aqueous iodine used to determine vitamin C content in sports drinks can be prepared by combining aqueous solutions of iodic acid (HIO_3) and hydroiodic acid (HI). (The products are aqueous iodine and liquid water.) Write a balanced equation for this reaction [⏮ Sample Problem 3.3]. (d) Write the net ionic equation for the reaction [⏮ Sample Problem 4.3]. (e) Determine the oxidation number for each element in the net ionic equation [⏮ Sample Problem 4.5].

New Pedagogy

Key Skills

Newly located immediately before the end-of-chapter problems, Key Skills pages are modules that provide a review of specific problem-solving techniques from that particular chapter. These are techniques the author knows are vital to success in later chapters. The Key Skills pages are designed to be easy for students to find touchstones to hone specific skills from earlier chapters—in the context of later chapters. The answers to the Key Skills Problems can be found in the Answer Appendix in the back of the book.



KEY SKILLS **Net Ionic Equations**

A molecular equation is necessary to do stoichiometric calculations [⏮ Section 3.3] but molecular equations often misrepresent the species in a solution.

Net ionic equations are preferable in many instances because they indicate more succinctly the species in solution and the actual chemical process that a chemical equation represents. Writing net ionic equations is an important part of solving a variety of problems including those involving precipitation reactions, redox reactions, and acid-base neutralization reactions. To write net ionic equations, you must draw on several skills from earlier chapters:

- Recognition of the common polyatomic ions [⏮ Section 2.6]
- Balancing chemical equations and labeling species with (*s*), (*l*), (*g*), or (*aq*) [⏮ Section 3.1]
- Identification of strong electrolytes, weak electrolytes, and nonelectrolytes [⏮ Section 4.1]

Writing a net ionic equation begins with writing and balancing the molecular equation. For example, consider the precipitation reaction that occurs when aqueous solutions of sodium iodide and lead(II) nitrate are combined.

$Pb(NO_3)_2(aq) + NaI(aq) \longrightarrow$

Exchanging the ions of the two aqueous reactants gives us the formulas of the products. The phases of the products are determined by considering the solubility guidelines [⏮ Tables 4.2 and 4.3].

$Pb(NO_3)_2(aq) + NaI(aq) \longrightarrow PbI_2(s) + NaNO_3(aq)$

We balance the equation and separate the soluble strong electrolytes to get the ionic equation.

$Pb(NO_3)_2(aq) + 2NaI(aq) \longrightarrow PbI_2(s) + 2NaNO_3(aq)$

$Pb^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + 2I^-(aq) \longrightarrow PbI_2(s) + 2Na^+(aq) + 2NO_3^-(aq)$

We then identify the spectator ions, those that are identical on both sides of the equation, and eliminate them.

$Pb^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + 2I^-(aq) \longrightarrow PbI_2(s) + 2Na^+(aq) + 2NO_3^-(aq)$

What remains is the net ionic equation.

$Pb^{2+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s)$

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Consider now the reaction that occurs when aqueous solutions of hydrochloric acid and potassium fluoride are combined.

$HCl(aq) + KF(aq) \longrightarrow$

Again, exchanging the ions of the two aqueous reactants gives us the formulas of the products.

$HCl(aq) + KF(aq) \longrightarrow HF(aq) + KCl(aq)$

This equation is already balanced. We separate soluble strong electrolytes into their constituent ions. In this case, although the products are both aqueous, only one is a strong electrolyte. The other, HF, is a *weak* electrolyte.

$H^+(aq) + Cl^-(aq) + K^+(aq) + F^-(aq) \longrightarrow HF(aq) + K^+(aq) + Cl^-(aq)$

We identify the spectator ions and eliminate them.

$H^+(aq) + Cl^-(aq) + K^+(aq) + F^-(aq) \longrightarrow HF(aq) + K^+(aq) + Cl^-(aq)$

What remains is the net ionic equation.

$H^+(aq) + F^-(aq) \longrightarrow HF(aq)$

You must be able to identify the species in solution as strong, weak, or nonelectrolytes so that you know which should be separated into ions and which should be left as molecular or formula units.

Key Skills Problems

4.1
What is the balanced net ionic equation for the precipitation of $FeSO_4(s)$ when aqueous solutions of K_2SO_4 and $FeCl_2$ are combined?

(a) $2K^+(aq) + SO_4^{2-}(aq) + Fe^{2+}(aq) + 2Cl^-(aq) \longrightarrow FeSO_4(s) + 2K^+(aq) + 2Cl^-(aq)$
(b) $Fe^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow FeSO_4(s)$
(c) $K_2SO_4(aq) + FeCl_2(aq) \longrightarrow FeSO_4(s) + 2KCl(aq)$
(d) $Fe^{2+}(aq) + 2SO_4^{2-}(aq) \longrightarrow FeSO_4(s)$
(e) $2K^+(aq) + SO_4^{2-}(aq) + Fe^{2+}(aq) + 2Cl^-(aq) \longrightarrow FeSO_4(s)$

4.2
Consider the following net ionic equation: $Cd^{2+}(aq) + 2OH^-(aq) \longrightarrow Cd(OH)_2(s)$. If the spectator ions in the ionic equation are $NO_3^-(aq) + K^+(aq)$, what is the molecular equation for this reaction?

(a) $CdNO_3(aq) + KOH(aq) \longrightarrow Cd(OH)_2(s) + KNO_3(aq)$
(b) $Cd^{2+}(aq) + NO_3^-(aq) + 2K^+(aq) + OH^-(aq) \longrightarrow Cd(OH)_2(s) + 2K^+(aq) + NO_3^-(aq)$
(c) $Cd(NO_3)_2(aq) + 2KOH(aq) \longrightarrow Cd(OH)_2(s) + 2KNO_3(aq)$
(d) $Cd(OH)_2(s) + 2KNO_3(aq) \longrightarrow Cd(NO_3)_2(aq) + 2KOH(aq)$
(e) $Cd^{2+}(aq) + NO_3^-(aq) + K^+(aq) + OH^-(aq) \longrightarrow Cd(OH)_2(s) + K^+(aq) + NO_3^-(aq)$

4.3
The net ionic equation for the neutralization of acetic acid ($HC_2H_3O_2$) with lithium hydroxide [$LiOH(aq)$] is

(a) $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$
(b) $H^+(aq) + C_2H_3O_2^-(aq) \longrightarrow HC_2H_3O_2(aq)$
(c) $HC_2H_3O_2(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$
(d) $HC_2H_3O_2(aq) + Li^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + LiC_2H_3O_2(aq)$
(e) $H^+(aq) + C_2H_3O_2^-(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$

4.4
When steel wool [$Fe(s)$] is placed in a solution of $CuSO_4(aq)$, the steel becomes coated with copper metal and the characteristic blue color of the solution fades. What is the net ionic equation for this reaction?

(a) $Fe(s) + CuSO_4(aq) \longrightarrow FeSO_4(aq) + Cu(s)$
(b) $Fe^{2+}(aq) + Cu(s) \longrightarrow Fe(s) + Cu^{2+}(aq)$
(c) $FeSO_4(aq) + Cu(s) \longrightarrow Fe(s) + CuSO_4(aq)$
(d) $Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$
(e) $Fe(s) + Cu(aq) \longrightarrow Fe(aq) + Cu(s)$

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New to the Fourth Edition

- **New End-of-Chapter Problems** have been added in response to user comments. These include additional conceptual problems, additional problems with limiting-reactant components, and updates of information in topical questions.
- **Key Skills** sections are newly located immediately before the end-of-chapter problems. These modules provide a review of specific problem-solving techniques that the author knows will be critical in later chapters. A unique approach, the context of these reviews combines that of the current chapter, and that of the later chapter(s) for which the specific skills will be important.
- **Continued development of truly comprehensive and consistent problem-solving.** Hundreds of worked examples (Sample Problems) help students get started learning how to approach and solve problems.
- **Updated Table of Contents** reflecting changes discussed in reviews and focus groups. The introduction of nomenclature has been reordered to put ionic compounds first—increasing the clarity of the subject for students.
- **SmartBook™ with Learning Resources.** Our adaptive SmartBook has been supplemented with additional learning resources tied to each learning objective to provide point-in-time help to students who need it.
- **More consistent use of H_3O^+ to represent the hydronium ion.** In graphics where space constraints require use of H^+, students are alerted to it and are reminded that the two different representations refer to the same aqueous species.

New and updated chapter content includes:
Incorporation of essential information from student notes into the main flow of text in each chapter. The remaining student notes are designed to help students over a variety of stumbling blocks. They include timely warnings about common errors, reminders of important information from previous chapters, and general information that helps place the material in an easily understood context.

Chapter 1—Expanded coverage of the treatment of units that are raised to powers

Chapter 2—Reorganization of nomenclature coverage

Chapter 3—New limiting-reactant problems

Chapter 4—New end-of-chapter problems, including limiting-reactant problems

Chapter 6—New chapter opener

Chapter 8—New problems involving polar molecules and percent ionic character

Chapter 9—New introduction of organic bond-line structures

Chapter 11—New Checkpoint questions

Chapter 13—New conceptual end-of-chapter problems

Chapter 14—New highly visual molecular-level illustrations of the effects of reactant concentration and temperature on reaction rate

Chapter 15—New conceptual end-of-chapter problems

Chapter 16—Consistent use of H_3O^+ to represent the hydronium ion. In graphics where space constraints require use of H^+, students are alerted to it and are reminded that the two different representations refer to the same aqueous species.

Chapter 18—New chapter opener and new conceptual end-of-chapter problems

Student Resources

All students will have access to **chemistry animations** for the animated Visualizing Chemistry figures as well as other chemistry animations in Connect. Within the text, the animations are mapped to the appropriate content.

Students will have access to innovative applications of new educational technologies. Based on their instructor's choices, students will have access to electronic homework and guided practice through **Connect.** Available questions include a variety of conceptual, static and algorithmic content chosen by the instructors specifically for their students. Connect is also a portal for McGraw-Hill SmartBook®, an exciting adaptive reading experience that formulates an individualized learning path for each student through an easy, intuitive interface and real-time diagnostic exercises.

Additionally, students can purchase a Study Guide containing material to practice problem-solving skills and a Student Solution Manual that contains detailed solutions and explanations for the odd-numbered problems in the main text.

For me, this text will always remain a work in progress. I encourage you to contact me with any comments or questions.

Julia Burdge
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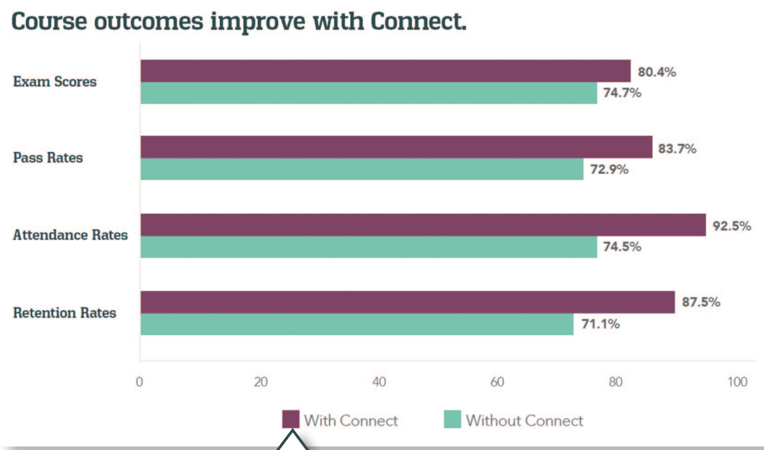

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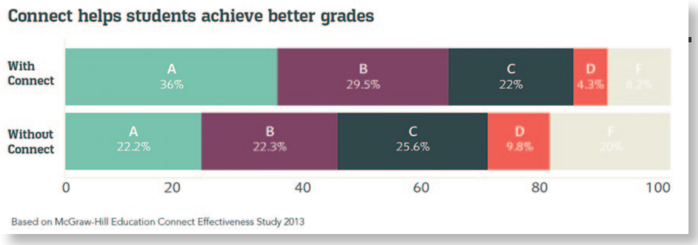

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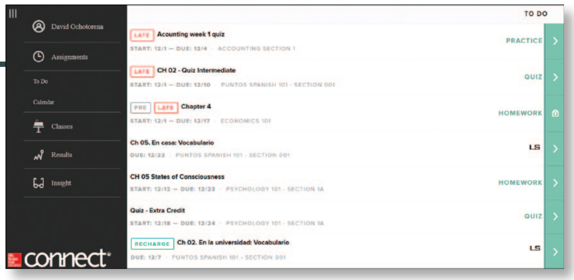

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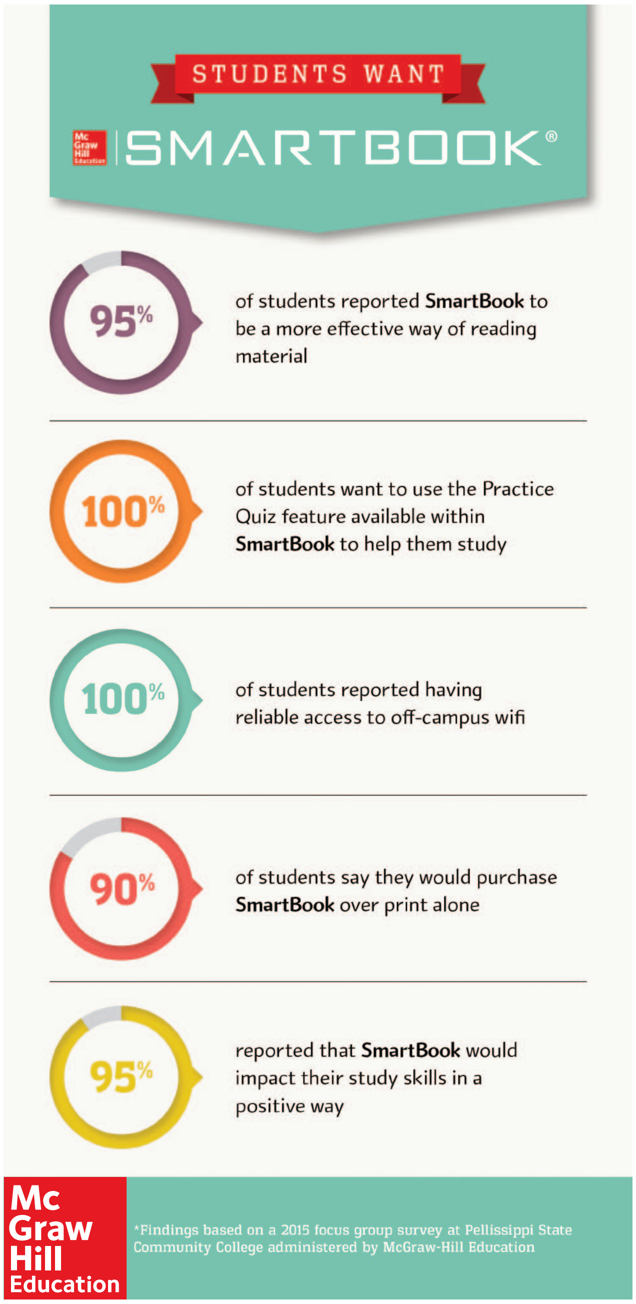

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My family, as always, continues to be there for me—no matter what.

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Preface
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Chemistry

CHAPTER 1

Chemistry: The Central Science

The "Epidemic Memorial" masks, on display at the Washington State History Museum in Tacoma, Washington, were created by five Native American artists. They represent the effects of smallpox and other diseases on the Native American population.

Credit: *Washington State History Research Center.*

In This Chapter, You Will Learn

Some of what chemistry is and how it is studied using the scientific method. You will learn about the system of units used by scientists and about expressing and dealing with the numbers that result from scientific measurements.

Before You Begin, Review These Skills

- Basic algebra
- Scientific notation [⏭ Appendix 1]

How the Scientific Method Helped Defeat Smallpox

To advance understanding of science, researchers use a set of guidelines known as the scientific method. The guidelines involve careful observations, educated reasoning, and the development of hypotheses and theories, which must undergo extensive testing. One of the most compelling examples of the success of the scientific method is the story of smallpox.

Smallpox is one of the diseases classified by the Centers for Disease Control and Prevention (CDC) as a Category A bioterrorism agent. This disease has had an immeasurable impact on human history. During the sixteenth century, European explorers brought smallpox with them to the Americas, devastating native populations and leaving them vulnerable to attack—in effect, shaping the conquest of the New World. In the twentieth century alone, the disease killed an estimated half a *billion* people worldwide—leaving many more permanently disfigured, blind, or both.

Student Note: *Category A* agents are those believed to pose the greatest potential threat to the public and that have a moderate to high potential for large-scale dissemination.

Late in the eighteenth century, an English doctor named Edward Jenner observed that even during outbreaks of smallpox in Europe, milkmaids seldom contracted the disease. He reasoned that when people who had frequent contact with cows contracted *cowpox,* a similar but far less harmful disease, they developed a natural immunity to smallpox. He predicted that intentional exposure to the cowpox virus would produce the same immunity. In 1796, Jenner exposed an 8-year-old boy named James Phipps to the cowpox virus using pus from the cowpox lesions of a milkmaid named Sarah Nelmes. Six weeks later, when Jenner then exposed Phipps to the smallpox virus, the boy did *not* contract the disease. Subsequent experiments using the same technique (later dubbed *vaccination* from the Latin *vacca* meaning "cow") confirmed that immunity to smallpox could be induced.

The last naturally occurring case of smallpox occurred in 1977 in Somalia. In 1980, the World Health Organization declared smallpox officially eradicated. This historic triumph over a dreadful disease, one of the greatest medical advances of the twentieth century, began with Jenner's astute observations, inductive reasoning, and careful experimentation—the essential elements of the *scientific method.*

Student Note: Although naturally occurring smallpox was wiped out worldwide, samples have been kept in research laboratories in the United States and the former Soviet Union, and several countries are now thought to have unauthorized stockpiles of the virus.

Until recently, almost everyone had a smallpox vaccine scar—usually on the upper arm.

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At the end of this chapter, you will be able to answer several questions related to the smallpox vaccine **[⏭ Applying What You've Learned, page 30]**.

1.1 The Study of Chemistry

Chemistry often is called the *central science* because knowledge of the principles of chemistry can facilitate understanding of other sciences, including physics, biology, geology, astronomy, oceanography, engineering, and medicine. ***Chemistry*** is the study of *matter* and the *changes* that matter undergoes. Matter is what makes up our bodies, our belongings, our physical environment, and in fact our universe. ***Matter*** is anything that has mass and occupies space.

Chemistry You May Already Know

You may already be familiar with some of the terms used in chemistry. Even if this is your first chemistry course, you may have heard of *molecules* and know them to be tiny pieces of a substance—much too tiny to see. Further, you may know that molecules are made up of *atoms,* even smaller pieces of matter. And even if you don't know what a chemical formula is, you probably know that H_2O is water. You may have used, or at least heard, the term *chemical reaction;* and you are undoubtedly familiar with a variety of chemical reactions, such as those shown in Figure 1.1.

The reactions in Figure 1.1 are all things that you can observe at the *macroscopic level*. In other words, these processes and their results are visible to the human eye. In studying chemistry, you will learn to visualize and understand these same processes at the *molecular level*.

Although it can take many different forms, all matter consists of various combinations of atoms of only a relatively small number of simple substances called *elements*. The properties of matter depend on which of these elements it contains and on how the atoms of those elements are arranged.

The Scientific Method

Experiments are the key to advancing our understanding of chemistry—or any science. Although not all scientists will necessarily take the same approach to experimentation, they all follow a set of guidelines known as the ***scientific method*** to add their results to the larger body of knowledge

(a)

(b)

(c)

(d)

(e)

Figure 1.1 Many familiar processes are chemical reactions: (a) The flame of a gas stove is the combustion of natural gas, which is primarily methane. (b) The bubbles produced when Alka-Seltzer dissolves in water are carbon dioxide, produced by a chemical reaction between two ingredients in the tablets. (c) The formation of rust is a chemical reaction that occurs when iron, water, and oxygen are all present. (d) Many baked goods "rise" as the result of a chemical reaction that produces carbon dioxide. (e) The glow produced when luminol is used to detect traces of blood in crime-scene investigations is the result of a chemical reaction.

Credit: *a: © Steve Allen/Getty Images; b: © McGraw-Hill Education/ Charles D. Winters, photographer; c: © Stockbyte/PunchStock; d: © Danilo Calilung/Corbis; e: © Jochen Tack/Alamy.*

FAQ What Do Molecules Look Like?

Molecules are far too small for us to observe them directly. An effective means of visualizing them is by the use of molecular models. Throughout this book, we will represent matter at the molecular level using *molecular art,* the two-dimensional equivalent of molecular models. In these pictures, atoms are represented as spheres, and atoms of particular elements are represented using specific colors. Table 1.1 lists some of the elements that you will encounter most often and the colors used to represent them in this book.

Molecular art can be of *ball-and-stick* models, in which the bonds connecting atoms appear as sticks [Figure 1.2(b)], or of *space-filling* models, in which the atoms appear to overlap one another [Figure 1.2(c)]. Ball-and-stick and space-filling models illustrate the specific, three-dimensional arrangement of the atoms. The ball-and-stick model does a good job of illustrating the arrangement of atoms, but exaggerates the distances between atoms, relative to their sizes. The space-filling model gives a more accurate picture of these *interatomic* distances but can obscure the details of the three-dimensional arrangement.

TABLE 1.1 Colors of Elements Commonly Used in Molecular Art

Hydrogen	Sodium
Boron	Phosphorus
Carbon	Sulfur
Nitrogen	Chlorine
Oxygen	Bromine
Fluorine	Iodine

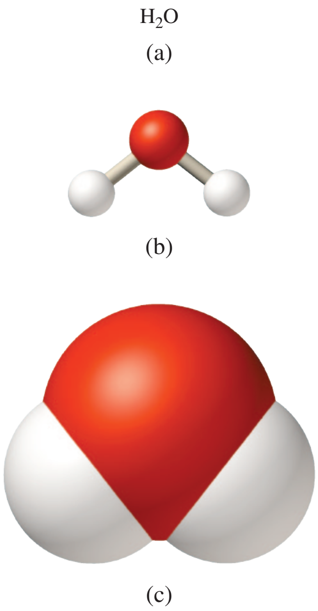


Figure 1.2 Water represented with a (a) molecular formula, (b) ball-and-stick model, and (c) space-filling model.

within a given field. The flowchart in Figure 1.3 illustrates this basic process. The method begins with the gathering of data via observations and experiments. Scientists study these data and try to identify *patterns* or *trends*. When they find a pattern or trend, they may summarize their findings with a ***law,*** a concise verbal or mathematical statement of a reliable relationship between phenomena. Scientists may then formulate a ***hypothesis,*** a tentative explanation for their observations. Further experiments are designed to test the hypothesis. If experiments indicate that the hypothesis is incorrect, the scientists go back to the drawing board, try to come up with a different

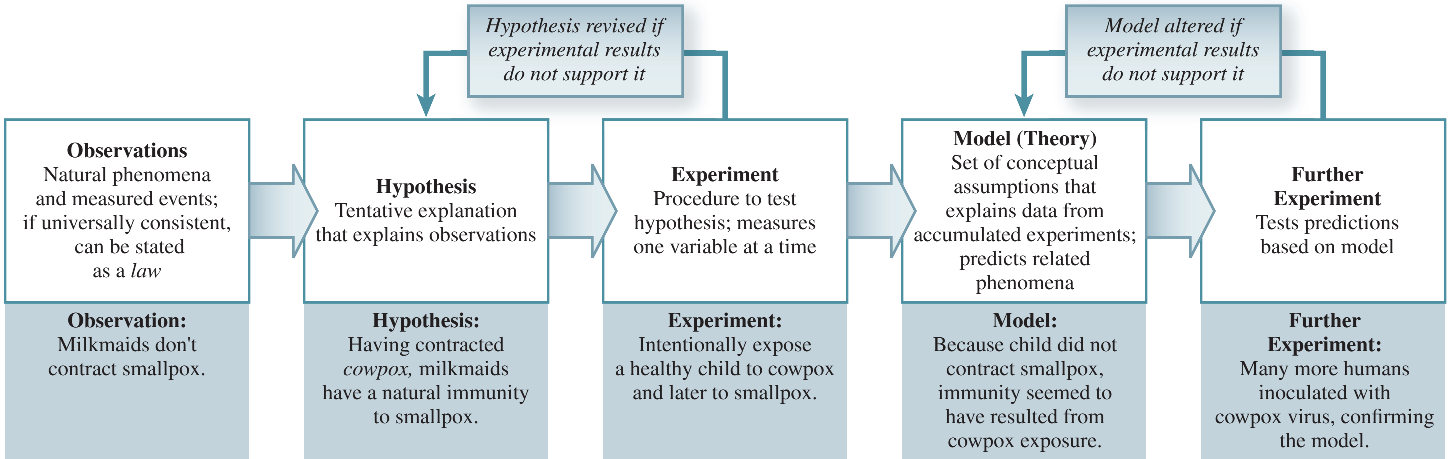


Figure 1.3 Flowchart of the scientific method.

interpretation of their data, and formulate a new hypothesis. The new hypothesis will then be tested by experiment. When a hypothesis stands the test of extensive experimentation, it may evolve into a theory. A ***theory*** is a unifying principle that explains a body of experimental observations and the laws that are based on them. Theories can also be used to predict related phenomena, so theories are constantly being tested. If a theory is disproved by experiment, then it must be discarded or modified so that it becomes consistent with experimental observations.

1.2 Classification of Matter

Student Note: Some books refer to *substances* as *pure substances*. These two terms generally mean the same thing although the adjective *pure* is unnecessary in this context because a substance is, by definition, pure.

Chemists classify matter as either a *substance* or a *mixture* of substances. A substance may be further categorized as either an *element* or a *compound*. A ***substance*** is a form of matter that has a definite (constant) composition and distinct properties. Examples are salt (sodium chloride), iron, water, mercury, carbon dioxide, and oxygen. Substances can be either elements (such as iron, mercury, and oxygen) or compounds (such as salt, water, and carbon dioxide). They differ from one another in composition and can be identified by appearance, smell, taste, and other properties.

States of Matter

All substances can, in principle, exist as a solid, a liquid, and a gas, the three physical states depicted in Figure 1.4. Solids and liquids sometimes are referred to collectively as the *condensed phases*. Liquids and gases sometimes are referred to collectively as *fluids*. In a solid, particles are held close together in an orderly fashion with little freedom of motion. As a result, a solid does not conform to the shape of its container. Particles in a liquid are close together but are not held rigidly in position; they are free to move past one another. Thus, a liquid conforms to the shape of the part of the container it fills. In a gas, the particles are separated by distances that are very large compared to the size of the particles. A sample of gas assumes both the shape and the volume of its container.

The three states of matter can be interconverted without changing the chemical composition of the substance. Upon heating, a solid (e.g., ice) will melt to form a liquid (water). Further heating will vaporize the liquid, converting it to a gas (water vapor). Conversely, cooling a gas will cause it to condense into a liquid. When the liquid is cooled further, it will freeze into the solid form. Figure 1.5 shows the three physical states of water.

Animation
Matter—three states of matter.

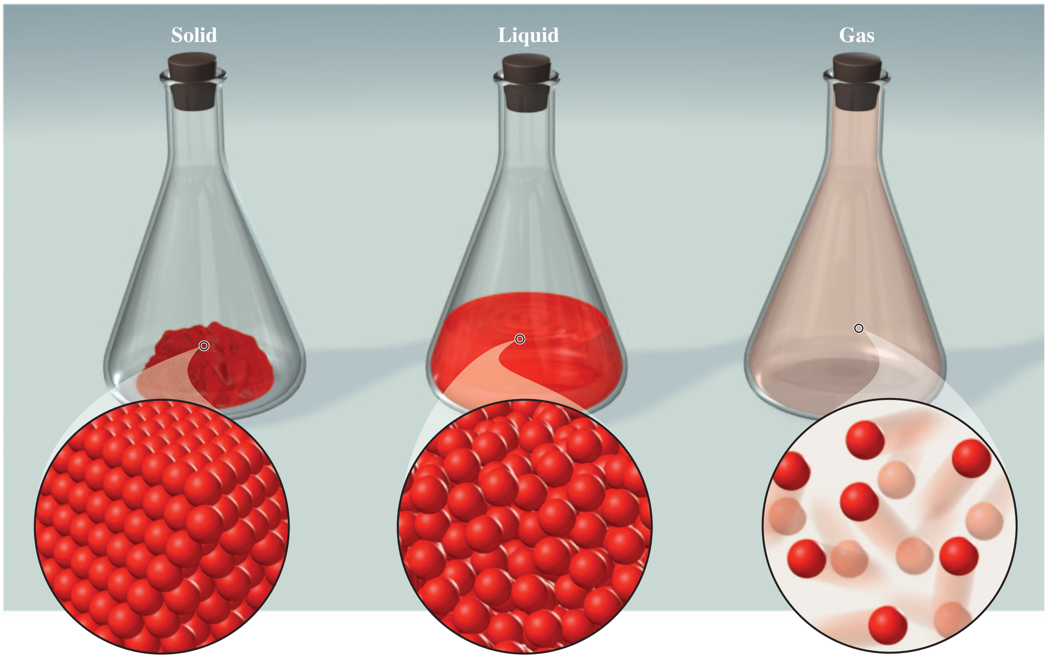


Figure 1.4 Molecular-level illustrations of a solid, liquid, and gas.

Figure 1.5 Water as a solid (ice), liquid, and gas. (We can't actually see water vapor, any more than we can see the nitrogen and oxygen that make up most of the air we breathe. When we see steam or clouds, what we are actually seeing is water vapor that has condensed upon encountering cold air.)

Credit: © *McGraw-Hill Education/Charles D. Winters, photographer.*

Elements

An ***element*** is a substance that cannot be separated into simpler substances by chemical means. Iron, mercury, oxygen, and hydrogen are just 4 of the 118 elements that have been identified. Most of the known elements occur naturally on Earth. The others have been produced by scientists via nuclear processes, which are discussed in Chapter 20. As shown in Figure 1.6(a) and (b), an element may consist of atoms or molecules.

For convenience, chemists use symbols of one or two letters to represent the elements. Only the first letter of an element's chemical symbol is capitalized. A list of the elements and their symbols appears on the inside front cover of this book. The symbols of some elements are derived from their Latin names—for example, Ag from *argentum* (silver), Pb from *plumbum* (lead), and Na from *natrium* (sodium)—while most of them come from their English names—for example, H for hydrogen, Co for cobalt, and Br for bromine.

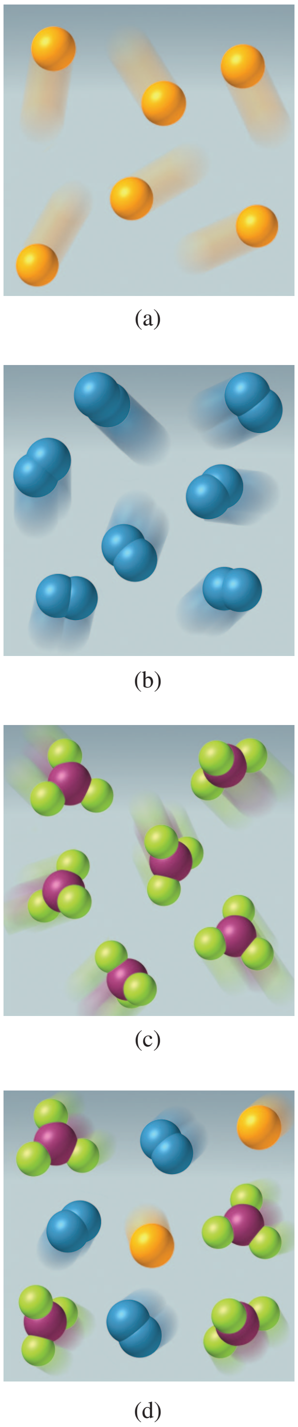


Figure 1.6 (a) Isolated atoms of an element. (b) Molecules of an element. (c) Molecules of a compound, consisting of more than one element. (d) A mixture of atoms of an element and molecules of an element and a compound.

Compounds

Most elements can combine with other elements to form compounds. Hydrogen gas, for example, burns in the presence of oxygen gas to form water, which has properties that are distinctly different from those of either hydrogen or oxygen. Thus, water is a ***compound,*** a substance composed of atoms of two or more elements chemically united in fixed proportions [Figure 1.6(c)]. The elements that make up a compound are called the compound's *constituent elements*. For example, the constituent elements of water are hydrogen and oxygen.

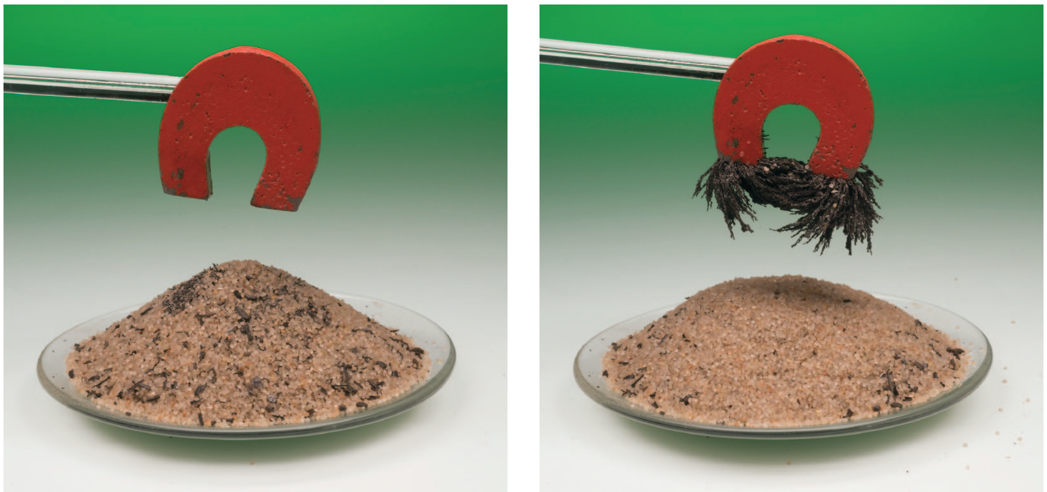
A compound cannot be separated into simpler substances by any physical process. (A physical process [⏭ Section 1.4] is one that does not change the identity of the matter. Examples of physical processes include boiling, freezing, and filtering.) Instead, the separation of a compound into its constituent elements requires a *chemical reaction*.

Student Note: A compound may consist of *molecules* or *ions*, which we will discuss in Chapter 2.

Mixtures

A ***mixture*** is a combination of two or more substances [Figure 1.6(d)] in which the substances retain their distinct identities. Like pure substances, mixtures can be solids, liquids, or gases. Some familiar examples are mixed nuts, 14-carat gold, apple juice, milk, and air. Mixtures do not have a universal constant composition. Therefore, samples of air collected in different locations will differ in composition because of differences in altitude, pollution, and other factors. Various brands of apple juice may differ in composition because of the use of different varieties of apples, or there may be differences in processing and packaging, and so on.

Mixtures are either *homogeneous* or *heterogeneous*. When we dissolve a teaspoon of sugar in a glass of water, we get a ***homogeneous mixture*** because the composition of the mixture is uniform throughout. If we mix sand with iron filings, however, the sand and the iron filings remain distinct and discernible from each other (Figure 1.7). This type of mixture is called a ***heterogeneous mixture*** because the composition is *not* uniform.



(a) (b)

Figure 1.7 (a) A heterogeneous mixture contains iron filings and sand. (b) A magnet is used to separate the iron filings from the mixture.

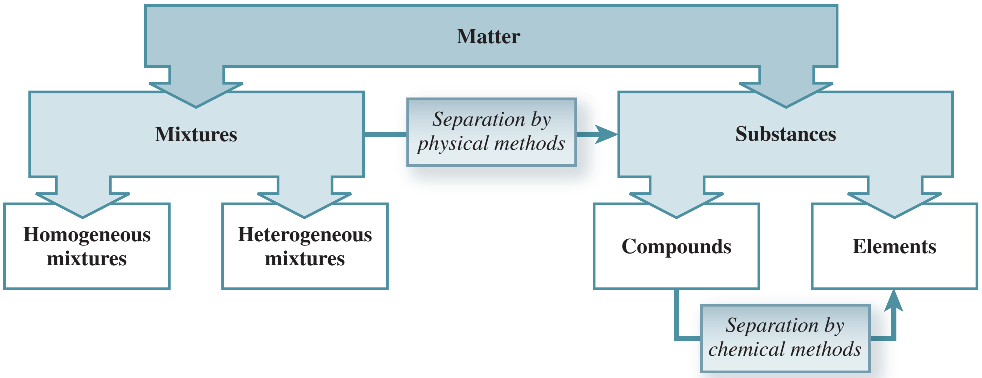


Figure 1.8 Flowchart for the classification of matter.

Mixtures, whether homogeneous or heterogeneous, can be separated by physical means into pure components without changing the identities of the components. Thus, sugar can be recovered from a water solution by evaporating the solution to dryness. Condensing the vapor will give us back the water component. To separate the sand–iron mixture, we can use a magnet to remove the iron filings from the sand, because sand is not attracted to the magnet [see Figure 1.7(b)]. After separation, the components of the mixture will have the same composition and properties as they did prior to being mixed. The relationships among substances, elements, compounds, and mixtures are summarized in Figure 1.8.

1.3 Scientific Measurement

Scientists use a variety of devices to measure the properties of matter. A meterstick is used to measure length; a burette, pipette, graduated cylinder, and volumetric flask are used to measure volume (Figure 1.9); a balance is used to measure mass; and a thermometer is used to measure temperature. Properties that can be measured are called *quantitative* properties because they are expressed using numbers. When we express a measured quantity with a number, though, we must always include the appropriate unit; otherwise, the measurement is meaningless. For example, to say that the depth of a swimming pool is 3 is insufficient to distinguish between one that is 3 *feet* (0.9 meter) and one that is 3 *meters* (9.8 feet) deep. Units are essential to reporting measurements correctly.

The two systems of units with which you are probably most familiar are the *English system* (foot, gallon, pound, etc.) and the *metric system* (meter, liter, kilogram, etc.). Although there has been an increase in the use of metric units in the United States in recent years, English units still are used commonly. For many years, scientists recorded measurements in metric units, but in 1960, the General Conference on Weights and Measures, the international authority on units, proposed a revised metric system for universal use by scientists. We will use both metric and revised metric (SI) units in this book.

Student Note: According to the U.S. Metric Association (USMA), the United States is "the only significant holdout" with regard to adoption of the metric system. The other countries that continue to use traditional units are Myanmar (formerly Burma) and Liberia.

SI Base Units

The revised metric system is called the ***International System of Units*** (abbreviated SI, from the French *Système Internationale d'Unités*). Table 1.2 lists the seven SI base units. All other units of measurement can be derived from these base units. The ***SI unit*** for *volume,* for instance, is derived by cubing the SI base unit for *length*. The prefixes listed in Table 1.3 are used to denote decimal fractions and multiples of SI units. This enables scientists to tailor the magnitude of a unit to a particular application. For example, the meter (m) is appropriate for describing the dimensions of a classroom, but the kilometer (km), 1000 m, is more appropriate for describing the distance between two cities. Units that you will encounter frequently in the study of chemistry include those for mass, temperature, volume, and density.

Mass

Although the terms *mass* and *weight* often are used interchangeably, they do not mean the same thing. Strictly speaking, weight is the force exerted by an object or sample due to gravity.

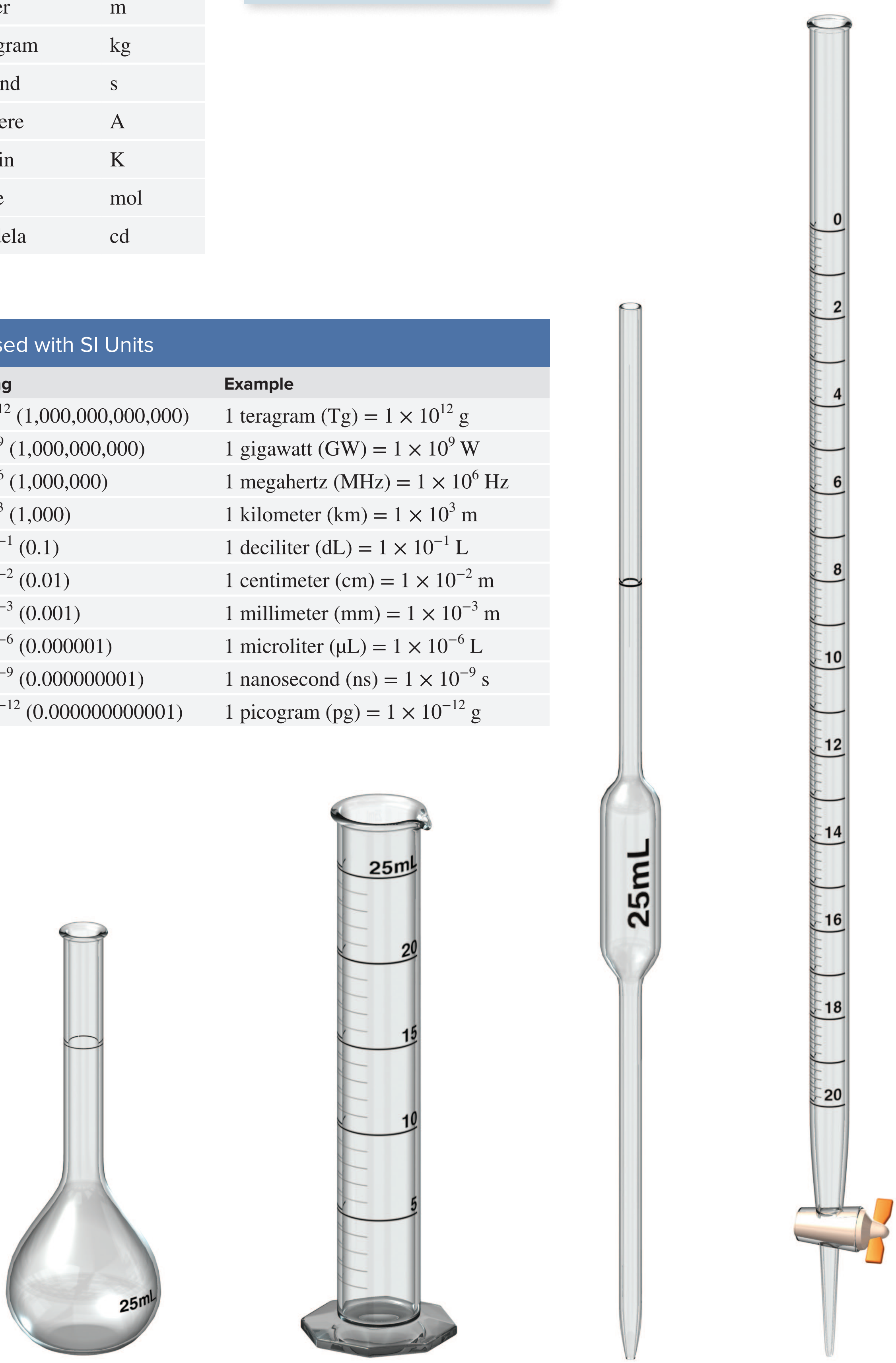
TABLE 1.2 Base SI Units

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

Student Note: Only one of the seven SI base units, the kilogram, itself contains a prefix.

TABLE 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
Tera-	T	1×10^{12} (1,000,000,000,000)	1 teragram (Tg) = 1×10^{12} g
Giga-	G	1×10^{9} (1,000,000,000)	1 gigawatt (GW) = 1×10^{9} W
Mega-	M	1×10^{6} (1,000,000)	1 megahertz (MHz) = 1×10^{6} Hz
Kilo-	k	1×10^{3} (1,000)	1 kilometer (km) = 1×10^{3} m
Deci-	d	1×10^{-1} (0.1)	1 deciliter (dL) = 1×10^{-1} L
Centi-	c	1×10^{-2} (0.01)	1 centimeter (cm) = 1×10^{-2} m
Milli-	m	1×10^{-3} (0.001)	1 millimeter (mm) = 1×10^{-3} m
Micro-	µ	1×10^{-6} (0.000001)	1 microliter (µL) = 1×10^{-6} L
Nano-	n	1×10^{-9} (0.000000001)	1 nanosecond (ns) = 1×10^{-9} s
Pico-	p	1×10^{-12} (0.000000000001)	1 picogram (pg) = 1×10^{-12} g



Volumetric flask
(a)

Graduated cylinder
(b)

Pipette
(c)

Burette
(d)

Figure 1.9 (a) A volumetric flask is used to prepare a precise volume of a solution for use in the laboratory. (b) A graduated cylinder is used to measure a volume of liquid. It is less precise than the volumetric flask. (c) A volumetric pipette is used to deliver a precise amount of liquid. (d) A burette is used to measure the volume of a liquid that has been added to a container. A reading is taken before and after the liquid is delivered, and the volume delivered is determined by subtracting the first reading from the second.