General, Organic, & Biological CHEMISTRY

Fifth Edition

Janice Gorzynski Smith

University of Hawaiʻi at Mānoa

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About the Author

Daniel C. Smith

Janice Gorzynski Smith was born in Schenectady, New York. She received an A.B. degree *summa cum laude* in chemistry at Cornell University and a Ph.D. in Organic Chemistry from Harvard University under the direction of Nobel Laureate E. J. Corey. During her tenure with the Corey group, she completed the total synthesis of the plant growth hormone gibberellic acid.

Following her postdoctoral work, Jan joined the faculty of Mount Holyoke College where she was employed for 21 years. During this time she was active in teaching chemistry lecture and lab courses, conducting a research program in organic synthesis, and serving as department chair. Her organic chemistry class was named one of Mount Holyoke's "Don't-miss courses" in a survey by *Boston* magazine. After spending two sabbaticals amidst the natural beauty and diversity in Hawai'i in the 1990s, Jan and her family moved there permanently in 2000. Most recently, she has served as a faculty member at the University of Hawai'i at Mānoa. In 2003, she received the Chancellor's Citation for Meritorious Teaching.

Jan resides in Hawai'i with her husband Dan, an emergency medicine physician, pictured with her hiking in Laos in 2019. She has four children and nine grandchildren. When not teaching, writing, or enjoying her family, Jan bikes, hikes, snorkels, and scuba dives in sunny Hawai'i, and time permitting, enjoys travel and Hawaiian quilting.

Dedicated to my family, especially Max, Oliver, Alijah, Koa, Logan, Elliott, Penelope, Otis, and Isabelle

About the cover Giant kelp, a type of marine algae that grows in dense forests in cold ocean waters, is a source of atmospheric chloromethane (CH_3Cl), a simple organic compound that contains the halogen chlorine. Chloromethane, a colorless gas with a faint odor, is also formed in forests by wood-rotting fungi and is released during volcanic eruptions. Because it is a key compound in the manufacture of polymers and drugs, chloromethane is extensively produced by the chemical industry, but most of the chloromethane in the atmosphere is natural in origin. In *General, Organic, & Biological Chemistry,* we learn about the chemical properties of compounds like chloromethane.

Brief Contents

Contents

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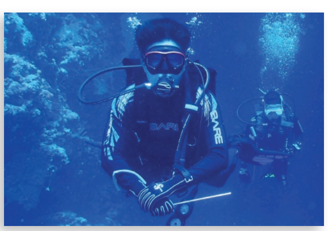

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20 Carbohydrates 743

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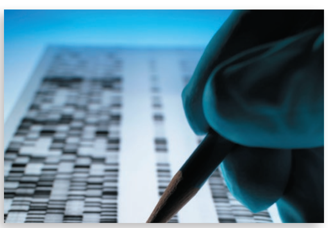
21 Amino Acids, Proteins, and Enzymes 787



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Preface

My goal in writing this text was to relate the fundamental concepts of general, organic, and biological chemistry to the world around us, and in this way illustrate how chemistry explains many aspects of everyday life. A key feature is the use of molecular art to illustrate and explain common phenomena we encounter every day. Each topic is broken down into small chunks of information that are more manageable and easily learned. Students are given enough detail to understand basic concepts, such as how soap cleans away dirt and why trans fats are undesirable in the diet, without being overwhelmed.

This textbook is written for students who have an interest in nursing, nutrition, environmental science, food science, and a wide variety of other health-related professions. The content of this book is designed for an introductory chemistry course with no chemistry prerequisite, and is suitable for either a two-semester sequence or a one-semester course. I have found that by introducing one new concept at a time, keeping the basic themes in focus, and breaking down complex problems into small pieces, many students in these chemistry courses acquire a new appreciation of both the human body and the larger world around them.

The Learning System Used in *General, Organic, & Biological Chemistry*

- **Writing Style** A concise writing style allows students to focus on learning major concepts and themes of general, organic, and biological chemistry. Relevant materials from everyday life are used to illustrate concepts, and topics are broken into small chunks of information that are more easily learned.
- **Chapter Outline and Chapter Goals** The chapter outline lists the main headings of the chapter to help students map out the organization of each chapter's content, and the chapter goals identify the key concepts that students will learn.
- **Chapter Review** The end-of-chapter summary sections are divided into parts: **Key Terms, Key Concepts, Key Equations, Key Reactions,** and **Key Skills,** with structures and examples to illustrate important concepts and skills.
- **Macro-to-Micro Illustrations** Because today's students are visual learners, and because visualizing molecular-level representations of macroscopic phenomena is critical to the understanding of any chemistry course, many illustrations in this text include photos or drawings of everyday objects, paired with their molecular representation, to help students understand the chemistry behind ordinary occurrences.
- **Problem Solving** Sample Problems lead students through the thought process tied to successful problem solving by employing Analysis and Solution parts. Sample Problems are paired with Practice Problems to allow students to apply what they have just learned. The Practice Problems are followed by More Practice lists to point students to end-of-chapter Problems that are similar in concept. Other Problems within the chapter build on the concepts learned in the Sample and Practice Problems. Students can verify their answers to all Practice Problems and odd-numbered in-chapter and end-of-chapter Problems at the end of each chapter.
- ***How To*'s** Key processes are taught to students in a straightforward and easy-to-understand manner by using examples and multiple, detailed steps to solving problems.
- **Applications** Common applications of chemistry to everyday life are found in margin-placed Health Notes, Consumer Notes, and Environmental Notes, as well as sections titled "Focus on Health & Medicine," "Focus on the Environment," and "Focus on the Human Body."

New To This Edition

General

CHAPTER REVIEW Chapter Review, which replaces Chapter Highlights at the end of each chapter, consists of Key Terms that are defined in the Glossary, Key Concepts, Key Equations, Key Reactions, and Key Skills. The Key Concepts and Key Skills sections use art and chemical structures to more clearly explain the key features detailed within the chapter. Key Skills, which presents the steps needed to solve important topics within the chapter, should be especially valuable for students learning stepwise processes.

STUDY TIPS Brief Study Tips have been added to the margins in Chapters 1, 4, 7, 11, 13, and 19 to help students develop general methods for solving recurrent types of problems, such as those that require a specific equation or drawing the products of an organic reaction.

PHOTOS Over one-half of the chapter-opening photos have been replaced with photos emphasizing relevant material within the chapter. More marginal photos of applications on topics such as glaucoma medications, plant-based burgers, and cannabis have also been added.

ART The colors in artwork throughout the text were revised for emphasis, clarity, and consistency.

PROBLEMS Over 150 new problems have been added.

Other New Coverage

Some of the new material added within specific chapters is listed below.

- Coverage on using a scientific calculator with scientific notation and logarithms has been expanded in Sections 1.6B and 9.6B. Tables with art that indicates what buttons should be pressed and what calculator displays will be shown are given.
- Figure 3.7 presents a succinct graphic on how to name different kinds of ionic compounds.
- Chapter 4 opens by presenting a new current topic, the effect of sunscreens like oxybenzone on the bleaching of coral reefs. Later in the chapter, a practice problem asks students to determine properties of oxybenzone from its ball-and-stick model. Recent research on oxybenzone is also discussed in Section 13.11.
- In Chapter 5, a new Table 5.4 has been added to summarize reaction types with both general cases and specific examples. More color has been added to both reactions and mathematical equations throughout this chapter to highlight the key components.
- The discussion of dialysis and osmosis in Section 8.9 has been edited to emphasize the distinction between these related concepts. Three new problems on this subject have been added.
- New material on using PET scans to visualize the brain in Alzheimer's patients has been added to Section 10.5.
- New material on anesthetics has been added to Section 11.6.
- New material on sources of methane in the atmosphere has been added to Section 12.6.
- The treatment of CFCs in Chapter 14 has been updated to include new compounds used as refrigerants that cause less harm to the ozone layer.
- Section 17.4B now contains information on prostaglandin analogues that are used in the treatment of glaucoma. Additional material is also added to Section 19.11A.
- Figure 19.1 now presents data on saturated fats, unsaturated oils, and trans fats in bar graph form for easier visualization of lipid content.
- The discussion of the HIV drug amprenavir in Section 21.11 has been updated.
- The art used in Figures 22.2, 22.3, 22.5, and 22.7 has been updated to provide better clarity on the complex biochemical processes shown.
- A brief discussion of COVID-19 has been added to Section 22.11.
- The art in many of the metabolic pathways in Chapters 23 and 24 has been edited to better illustrate the key features of the steps in these schemes.

Our Commitment to Serving Teachers and Learners

TO THE INSTRUCTOR Writing a chemistry textbook is a colossal task. Teaching chemistry for over 30 years at both a private, liberal arts college and a large state university has given me a unique perspective with which to write this text. I have found that students arrive with vastly different levels of preparation and widely different expectations for their college experience. As an instructor and now an author, I have tried to channel my love and knowledge of chemistry into a form that allows this spectrum of students to understand chemical science more clearly, and then see everyday phenomena in a new light.

TO THE STUDENT I hope that this text and its ancillary program will help you to better understand and appreciate the world of chemistry. My interactions with thousands of students in my long teaching career have profoundly affected the way I teach and write about chemistry, so please feel free to email me with any comments or questions at jgsmith@hawaii.edu.

Learning Resources for Instructors and Students

Student Solutions Manual

Each chapter contains the solutions to all in-chapter problems, as well as the solutions to all odd-numbered end-of-chapter problems.

Instructor's Solutions Manual

This supplement contains complete, worked-out solutions for all the end-of-chapter problems in the text. It can be found within the Instructor's Resources on Connect's online learning center.

Presentation Tools

Accessed from your textbook's Instructor's Resources, **Presentation Tools** is an online digital library containing photos and artwork that can be used to create customized lectures, visually enhanced tests and quizzes, compelling course websites, or attractive printed support materials. All assets are copyrighted by McGraw-Hill Higher Education but can be used by instructors for classroom purposes. The visual resources in this collection also include:

- **PowerPoint Slides:** For instructors who prefer to create their lectures from scratch, all illustrations, photos, and tables are pre-inserted by chapter into PowerPoint slides.
- **Accessible Lecture Outlines:** Lecture notes incorporating illustrations have been written for the fifth edition text. They are provided in PowerPoint format so that you may use these lectures as written or customize them to fit your lecture.

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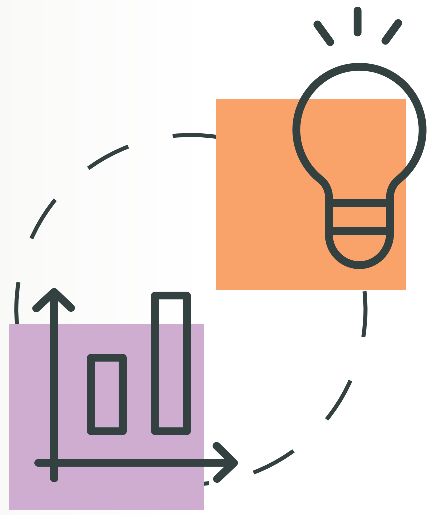

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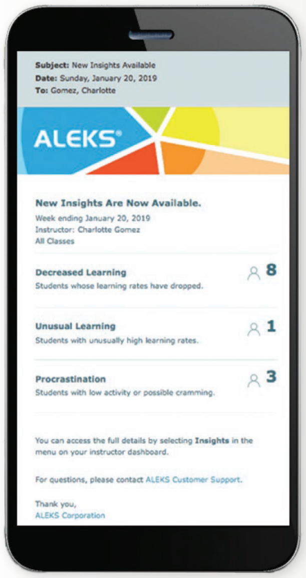

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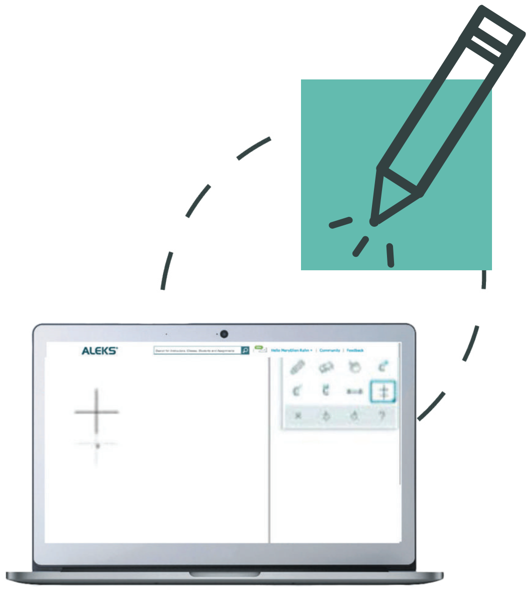

Engage Students Outside the Classroom with the MolecEd Organic Drawing Tool

Enhance student's learning experience with added interactivity for drawing Haworth and Fisher Projections. The MolecEd Organic drawing tool mirrors the drawing actions of paper and pencil and allows you to deepen critical thinking skills of students with assignments for practice and application.

SUPPORT AT every step

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Spend less time troubleshooting and more time fostering breakthrough moments with the help of ALEKS's dedicated, best-in-class implementation and tech support teams, its 99% uptime, and our collection of tools and resources designed to support you before, during, and beyond the first day of class.

*visit **http://bit.ly/SupportAtEveryStep** for real-world support and guidance

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Winner of 2019 Digital Edge 50 Award for Data Analytics!

Acknowledgments

With each new edition published, I am incredibly thankful to work with a group of knowledgeable and hard-working publishing professionals at McGraw-Hill.

I express deep gratitude to Senior Product Developer Mary Hurley who has become a dear friend as we have worked together in the successful publication of many editions. Special thanks also go to freelance Developmental Editor John Murdzek, whose meticulous attention to detail and breadth of knowledge in chemistry enhance each new edition. Thanks are also due to Portfolio Manager Michelle Hentz and the production, marketing, and sales teams for their ability to create such stunning texts and bring them to an ever-widening audience.

Finally, I thank my family for their support and patience during the long process of publishing a textbook. My husband Dan, an emergency medicine physician who continues to work in far-flung places (Cambodia, Bhutan, Fiji, and the Solomon Islands), took several photos that appear in the text and served as a consultant for many medical applications.

REVIEWERS

The following people were instrumental in reading and providing feedback on previous editions of the text, which helped to shape my ideas into cohesive pages:

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Matter and Measurement

1

Determining the weight and length of a newborn are common measurements performed by healthcare professionals.
Daniel C. Smith

CHAPTER OUTLINE

CHAPTER GOALS

In this chapter you will learn how to:

1. Describe the three states of matter
2. Classify matter as a pure substance, mixture, element, or compound
3. Report measurements using the metric units of length, mass, and volume
4. Use significant figures
5. Use scientific notation for very large and very small numbers
6. Use conversion factors to convert one unit to another
7. Convert temperature from one scale to another
8. Define density and specific gravity and use density to calculate the mass or volume of a substance

Why Study ...

Matter and Measurement?

Everything you touch, feel, or taste is composed of chemicals—that is, **matter**—so an understanding of its composition and properties is crucial to our appreciation of the world around us. Some matter—lakes, trees, sand, and soil—is naturally occurring, whereas other examples of matter—aspirin, CDs, nylon fabric, plastic syringes, and vaccines—are made by humans. To understand the properties of matter, as well as how one form of matter is converted to another, we must also learn about measurements. Following a recipe, pumping gasoline, and figuring out drug dosages involve manipulating numbers. Thus, Chapter 1 begins our study of chemistry by examining the key concepts of matter and measurement.

1.1 Chemistry—The Science of Everyday Experience

What activities might occupy the day of a typical student? You may have done some or all of the following tasks: eaten some meals, drunk coffee or cola, taken a shower with soap, checked email on a computer, ridden a bike or car to a part-time job, taken an aspirin to relieve a headache, and spent some of the evening having snacks and refreshments with friends. Perhaps, without your awareness, your life was touched by chemistry in each of these activities. What, then, is this discipline we call **chemistry?**

- *Chemistry* is the study of matter—its composition, properties, and transformations.

What is **matter?**

- *Matter* is anything that has mass and takes up volume.

In other words, **chemistry studies anything that we touch, feel, see, smell, or taste,** from simple substances like water and salt, to complex substances like proteins and carbohydrates that combine to form the human body. Some matter—cotton, sand, an apple, and the cardiac drug digoxin—is **naturally occurring,** meaning it is isolated from natural sources. Other substances—nylon, Styrofoam, the plastic used in soft drink bottles, and the pain reliever ibuprofen—are **synthetic,** meaning they are produced by chemists in the laboratory (Figure 1.1).

Figure 1.1 Naturally Occurring and Synthetic Materials

a. Naturally occurring materials

b. Synthetic materials

Matter occurs in nature or is synthesized in the lab. (a) Sand and apples are two examples of natural materials. Cotton fabric is woven from cotton fiber, obtained from the cotton plant. The drug digoxin (trade name Lanoxin), widely prescribed for decades for patients with congestive heart failure, is extracted from the leaves of the woolly foxglove plant. (b) Nylon was the first synthetic fiber made in the laboratory. It quickly replaced the natural fiber silk in parachutes and ladies' stockings. Styrofoam and PET (polyethylene terephthalate), the plastic used for soft drink bottles, are strong yet lightweight synthetic materials used for food storage. Over-the-counter pain relievers like ibuprofen are synthetic. The starting materials for all of these useful products are obtained from petroleum.
(a–b): Jill Braaten/McGraw-Hill

Figure 1.2 Transforming Natural Materials into Useful Synthetic Products

(a) Latex, the sticky liquid that oozes from a rubber tree when it is cut, is too soft for most applications. (b) Vulcanization converts latex into the stronger, elastic rubber used in tires and other products. *(a): Suphatthra China/Shutterstock; (b): Roy McMahon/Fuse/Getty Images*

Sometimes a chemist studies what a substance is made of, whereas at other times, the focus may be how to convert one material into a new material with unique and useful properties. As an example, naturally occurring rubber exists as the sticky liquid latex, which is too soft for most applications. The laboratory process of vulcanization converts it into the stronger, more elastic material used in tires and other products (Figure 1.2).

Chemistry is truly the science of everyday experience. Soaps and detergents, newspapers and CDs, condoms and oral contraceptives, Tylenol and penicillin—all of these items are products of chemistry. Without a doubt, advances in chemistry have transformed life in modern times.

Problem 1.1

Imagine that your job as a healthcare professional is to take a blood sample from a patient and store it in a small container in a refrigerator until it is picked up for analysis in the hospital lab. You might have to put on gloves and a mask, use a plastic syringe with a metal needle, store the sample in a test tube or vial, and place it in a cold refrigerator. Pick five objects you might encounter during the process and decide if they are made of naturally occurring or synthetic materials.

1.2 States of Matter

Matter exists in three common states—solid, liquid, and gas.

- A *solid* has a definite volume, and maintains its shape regardless of the container in which it is placed. The particles of a solid lie close together, and are arranged in a regular three-dimensional array.
- A *liquid* has a definite volume, but takes on the shape of the container it occupies. The particles of a liquid are close together, but they can randomly move around, sliding past one another.
- A *gas* has no definite shape or volume. The particles of a gas move randomly and are separated by a distance much larger than their size. The particles of a gas expand to fill the volume and assume the shape of whatever container they are put in.

For example, water exists in its solid state as ice or snow, liquid state as liquid water, and gaseous state as steam or water vapor. Blow-up circles like those in Figure 1.3 will be used commonly in this text to indicate the composition and state of the particles that compose a substance. In this molecular art, different types of particles are shown in color-coded spheres, and the distance between the spheres signals its state—solid, liquid, or gas.

Matter is characterized by its **physical properties** and **chemical properties.**

- *Physical properties* are those that can be observed or measured without changing the composition of the material.

Common physical properties include melting point (mp), boiling point (bp), solubility, color, and odor. **A *physical change* alters a substance without changing its composition.** The most common physical changes are **changes in state**—that is, the **conversion of matter from one state to another.** Melting an ice cube to form liquid water, and boiling liquid water to form steam, are two examples of physical changes. Water is the substance at the beginning and end of both physical changes. More details about physical changes are discussed in Chapter 7.

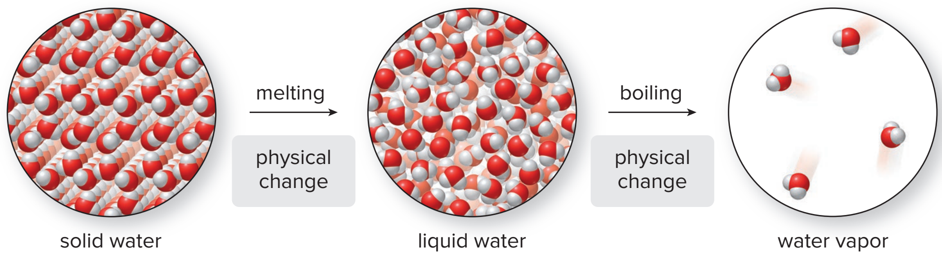
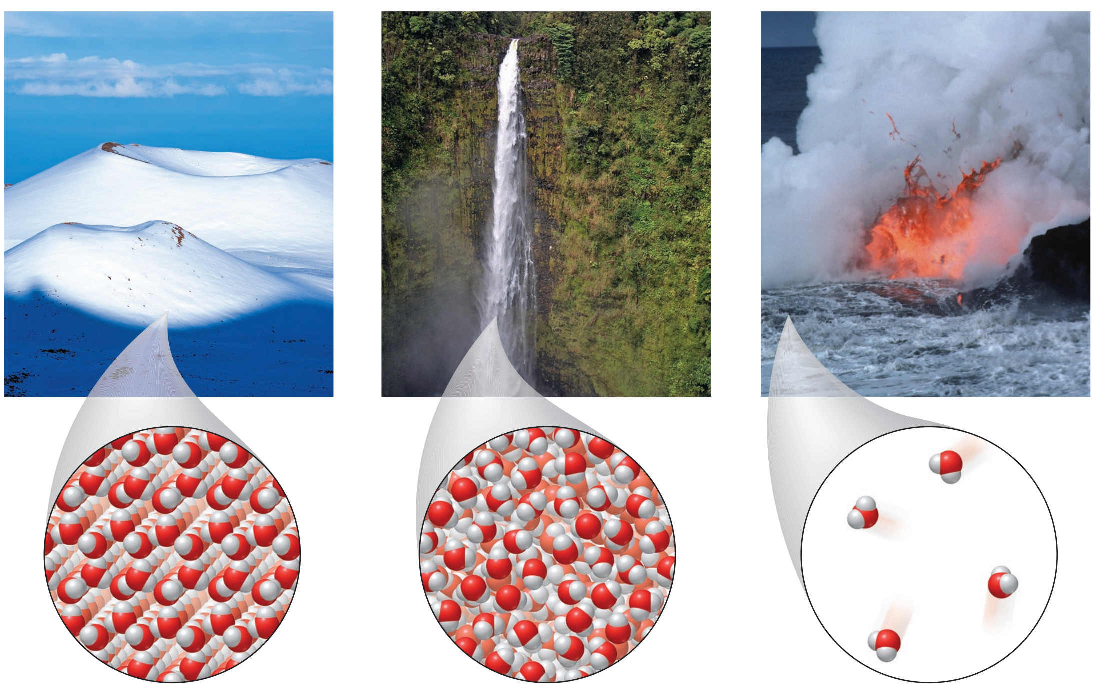


Figure 1.3 The Three States of Water—Solid, Liquid, and Gas

a. Solid water

b. Liquid water

c. Gaseous water

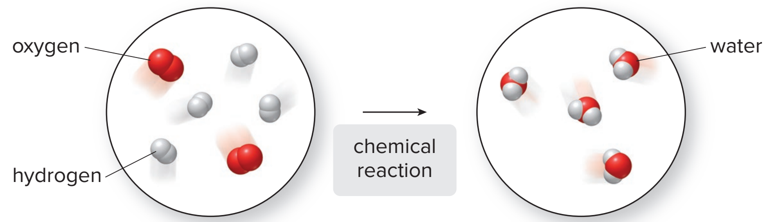


- The particles of a solid are **close together** and **highly organized.** (Photo: snow-capped Mauna Kea on the Big Island of Hawaii)
- The particles of a liquid are **close together** but **more disorganized** than the solid. (Photo: Akaka Falls on the Big Island of Hawaii)
- The particles of a gas are **far apart** and **disorganized.** (Photo: steam formed by a lava flow on the Big Island of Hawaii)

Each red sphere joined to two gray spheres represents a single water particle. In proceeding from left to right, from solid to liquid to gas, the molecular art shows that the level of organization of the water particles decreases. Color-coding and the identity of the spheres within the particles will be addressed in Chapter 2. *(a): Alvis Upitis/Getty Images; (b): Daniel C. Smith; (c): Source: T.J. Takahash/USGS*

- *Chemical properties* are those that determine how a substance can be converted to another substance.

A *chemical change,* or a *chemical reaction,* converts one material into another. The conversion of hydrogen and oxygen into water is a chemical reaction because the composition of the material is different at the beginning and end of the process. Chemical reactions are discussed in Chapters 5 and 6.

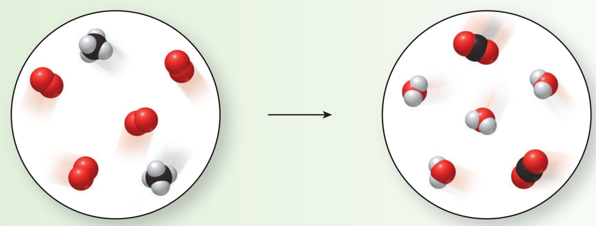


Problem 1.2

Characterize each process as a physical change or a chemical change: (a) making ice cubes; (b) burning natural gas; (c) silver jewelry tarnishing; (d) a pile of snow melting; (e) fermenting grapes to produce wine.

Problem 1.3

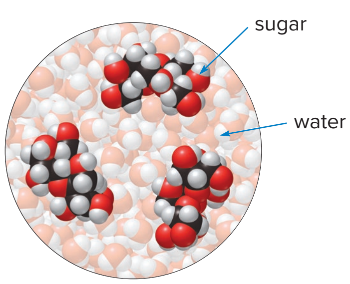
Does the molecular art represent a chemical change or a physical change? Explain your choice.



1.3 Classification of Matter

All matter can be classified as either a **pure substance** or a **mixture.**

- A *pure substance* is composed of a single component and has a constant composition, regardless of the sample size and the origin of the sample.



sugar dissolved in water

mixture

A pure substance, such as water or table sugar, can be characterized by its physical properties, because these properties do not change from sample to sample. **A pure substance cannot be broken down to other pure substances by any physical change.**

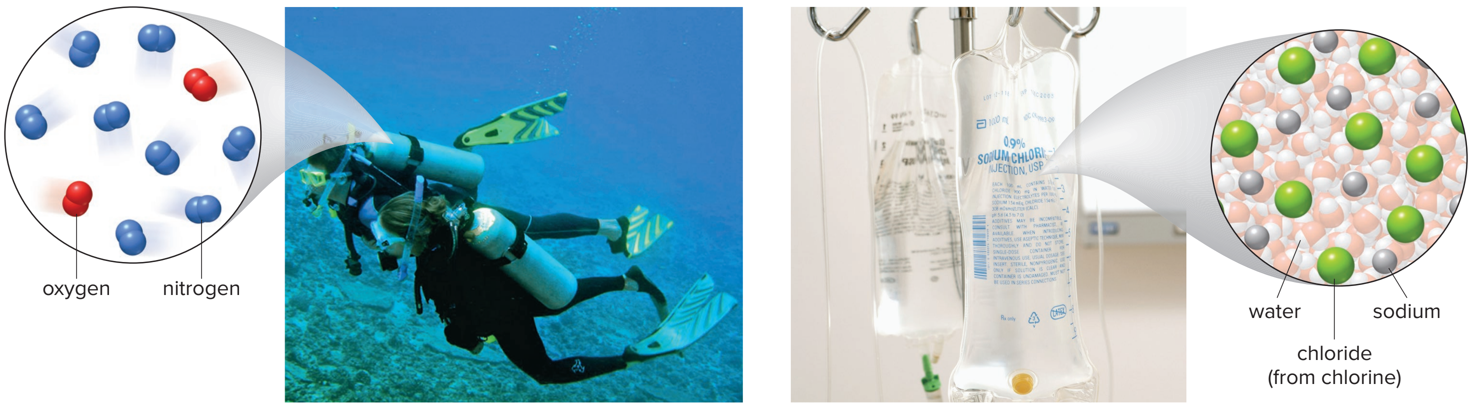
- A *mixture* is composed of more than one substance. The composition of a mixture can vary depending on the sample.

The physical properties of a mixture may also vary from one sample to another. **A mixture can be separated into its components by physical changes.** Dissolving table sugar in water forms a mixture, whose sweetness depends on the amount of sugar added. If the water is allowed to evaporate from the mixture, pure table sugar and pure water are obtained.

Figure 1.4 Two Examples of Mixtures

a. Two gases

b. A solid and a liquid



(a): Daniel C. Smith; (b): Janis Christie/Digital Vision/Alamy Stock Photo

Mixtures can be formed from solids, liquids, and gases, as shown in Figure 1.4. The compressed air breathed by a scuba diver consists mainly of the gases oxygen and nitrogen. A saline solution used in an IV bag contains solid sodium chloride (table salt) dissolved in liquid water.

A pure substance is classified as either an **element** or a **compound.**

- An *element* is a pure substance that cannot be broken down into simpler substances by a chemical reaction.
- A *compound* is a pure substance formed by chemically combining (joining together) two or more elements.

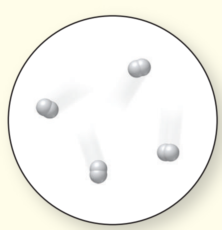
An alphabetical list of elements is located in Appendix A. The elements are commonly organized into a periodic table, shown in Appendix B, and discussed in much greater detail in Section 2.4.

Nitrogen gas, aluminum foil, and copper wire are all elements. Water is a compound because it is composed of the elements hydrogen and oxygen. Table salt, sodium chloride, is also a compound because it is formed from the elements sodium and chlorine (Figure 1.5). Although only 118 elements are currently known, over 50 million compounds occur naturally or have been synthesized in the laboratory. We will learn much more about elements and compounds in Chapter 2. Figure 1.6 summarizes the categories into which matter is classified.

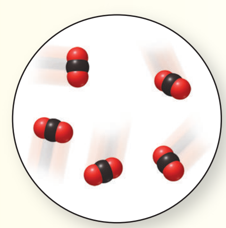
Sample Problem 1.1 Using Molecular Art for an Element and a Compound

Classify each example of molecular art as an element or a compound:

a.



b.



Analysis

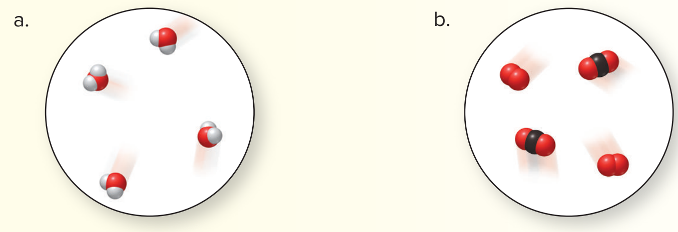
In molecular art, an element is composed of spheres of the same color, whereas a compound is composed of spheres of different colors.

Solution

Representation (a) is an element because each particle contains only gray spheres.
Representation (b) is a compound because each particle contains both red and black spheres.

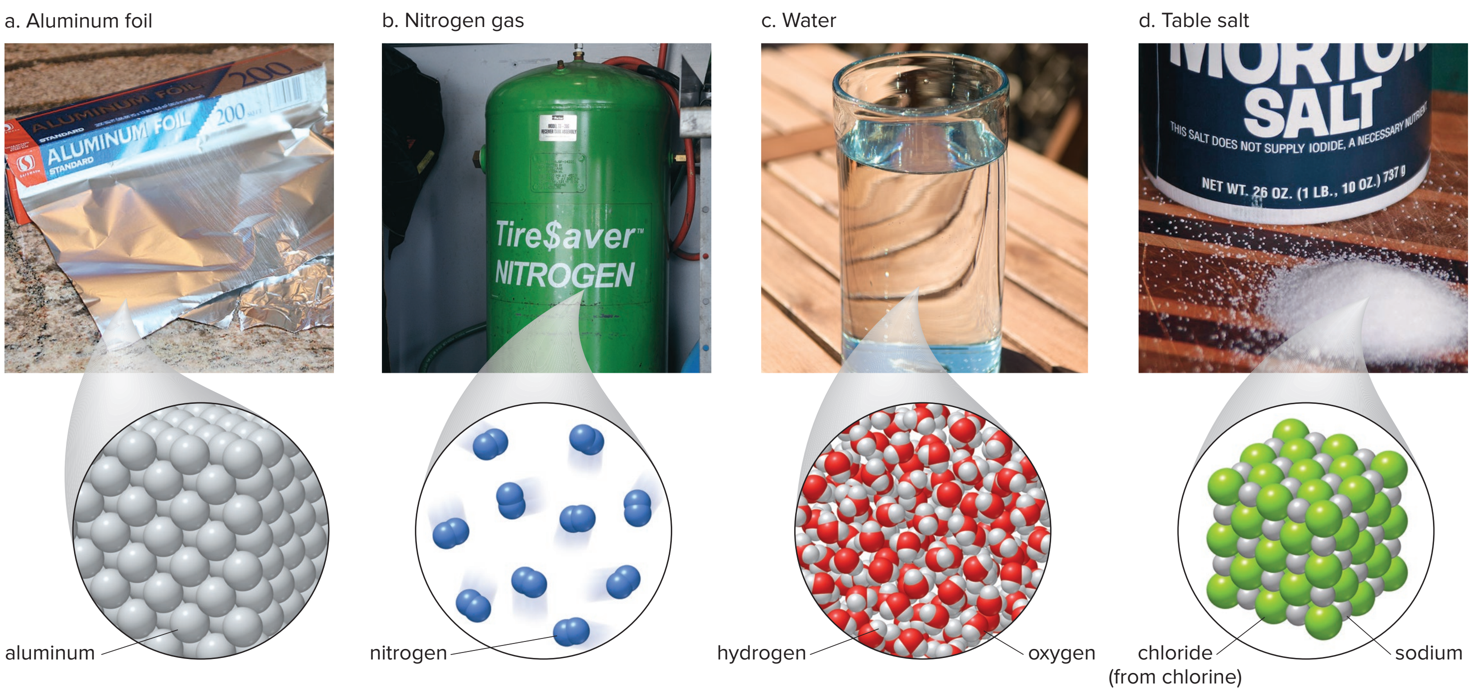
Practice Problem 1.1

Classify each example of molecular art as a pure substance or a mixture:



More Practice: Try Problems 1.19–1.21.

Figure 1.5 Elements and Compounds

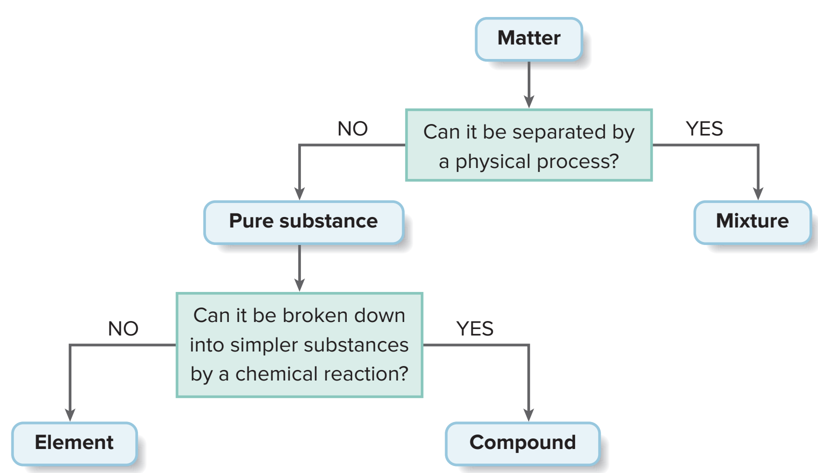


- Aluminum foil and nitrogen gas are elements. **The molecular art for an element shows spheres of *one* color only.** Thus, aluminum is a solid shown with gray spheres, whereas nitrogen is a gas shown with blue spheres. Water and table salt are compounds. Color-coding of the spheres used in the molecular art indicates that water is composed of two elements—hydrogen shown as gray spheres, and oxygen shown in red. Likewise, the gray (sodium) and green (chlorine) spheres illustrate that sodium chloride is formed from two elements as well.

(a): Daniel C. Smith; (b): Keith Eng, 2008; (c): Jill Braaten/McGraw-Hill; (d): Daniel C. Smith

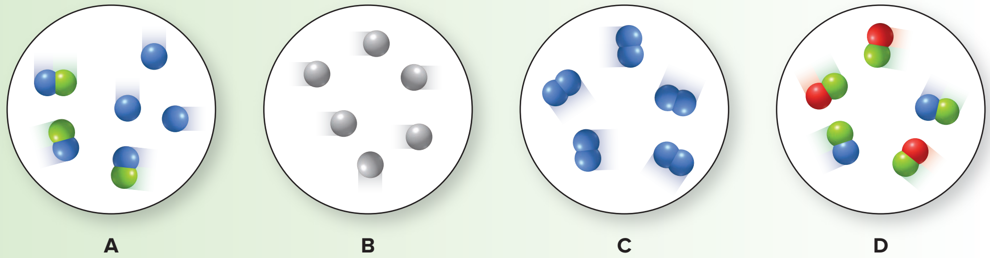
Figure 1.6

Classification of Matter



Problem 1.4

(a) Which representation(s) of molecular art illustrate pure elements? (b) Which representation(s) illustrate a mixture of a compound and an element? (c) Which representation(s) illustrate a mixture of two compounds?



Problem 1.5

Classify each item as a pure substance or a mixture: (a) blood; (b) ocean water; (c) a piece of wood; (d) a chunk of ice.

Problem 1.6

Classify each item as an element or a compound: (a) the gas inside a helium balloon; (b) table sugar; (c) the rust on an iron nail; (d) aspirin. All elements are listed alphabetically in Appendix A.

In 1960, the **International System of Units** was formally adopted as the uniform system of units for the sciences. **SI units,** as they are called, are based on the metric system, but the system recommends the use of some metric units over others. SI stands for the French words *Système Internationale.*

1.4 Measurement

Anytime you check your weight on a scale, measure the ingredients of a recipe, or figure out how far it is from one location to another, you are measuring a quantity. Measurements are routine for healthcare professionals who use weight, blood pressure, pulse, and temperature to chart a patient's progress.

Roy Hsu/Photographer's Choice RF/Getty Images

- **Every measurement is composed of a *number* and a *unit*.**

Reporting the value of a measurement is meaningless without its unit. For example, if you were told to give a patient an aspirin dosage of 325, does this mean 325 ounces, pounds, grams, milligrams, or tablets? Clearly there is a huge difference among these quantities.

1.4A The Metric System

In the United States, most measurements are made with the **English system,** using units like miles (mi), gallons (gal), pounds (lb), and so forth. A disadvantage of this system is that the units are not systematically related to each other and require memorization. For example, 1 lb = 16 oz, 1 gal = 4 qt, and 1 mi = 5,280 ft.

CONSUMER NOTE

The metric system is slowly gaining acceptance in the United States, as seen in the gallon jug of milk and the two-liter bottle of soda. *Jill Braaten*

Scientists, health professionals, and people in most other countries use the **metric system,** with units like meter (m) for length, gram (g) for mass, and liter (L) for volume. The metric system is slowly gaining popularity in the United States. The weight of packaged foods is often given in both ounces and grams. Distances on many road signs are shown in miles and kilometers. Most measurements in this text will be reported using the metric system, but learning to convert English units to metric units is also a necessary skill that will be illustrated in Section 1.7.

The important features of the metric system are the following:

- Each type of measurement has a base unit—the meter (m) for length; the gram (g) for mass; the liter (L) for volume; the second (s) for time.
- All other units are related to the base unit by powers of 10.
- The prefix of the unit name indicates if the unit is *larger* or *smaller* than the base unit.

The base units of the metric system are summarized in Table 1.1, and the most common prefixes used to convert the base units to smaller or larger units are summarized in Table 1.2. **The same prefixes are used for all types of measurement.** For example, the prefix *kilo-* means 1,000 times as large. Thus,

1 **kilo**meter = **1,000** meters	or	1 km = 1,000 m
1 **kilo**gram = **1,000** grams	or	1 kg = 1,000 g
1 **kilo**liter = **1,000** liters	or	1 kL = 1,000 L

The prefix *milli-* means one-thousandth as large (1/1,000 or 0.001). Thus,

1 **milli**meter = **0.001** meter	or	1 mm = 0.001 m
1 **milli**gram = **0.001** gram	or	1 mg = 0.001 g
1 **milli**liter = **0.001** liter	or	1 mL = 0.001 L

Table 1.1 Metric Units

Quantity	Metric Base Unit	Symbol
Length	Meter	m
Mass	Gram	g
Volume	Liter	L
Time	Second	s

The metric symbols are all lowercase except for the unit **liter** (L) and the prefixes **mega-** (M) and **giga-** (G).

Table 1.2 Common Prefixes Used for Metric Units

Prefix	Symbol	Meaning	Numerical Value[a]	Scientific Notation[b]
Giga-	G	Billion	1,000,000,000.	10^{9}
Mega-	M	Million	1,000,000.	10^{6}
Kilo-	k	Thousand	1,000.	10^{3}
Deci-	d	Tenth	0.1	10^{-1}
Centi-	c	Hundredth	0.01	10^{-2}
Milli-	m	Thousandth	0.001	10^{-3}
Micro-	µ[c]	Millionth	0.000 001	10^{-6}
Nano-	n	Billionth	0.000 000 001	10^{-9}

[a]Numbers that contain five or more digits to the right of the decimal point are written with a small space separating each group of three digits.
[b]How to express numbers in scientific notation is explained in Section 1.6.
[c]The symbol µ is the lowercase Greek letter mu. The prefix *micro-* is sometimes abbreviated as **mc.**

Sample Problem 1.2 Translating a Measurement into a Metric Unit

What term is used for each of the following units: (a) a thousand seconds; (b) a millionth of a gram?

Analysis

Use Table 1.2 to determine the prefix, and add the prefix before the base unit.

Solution

a. The prefix ***kilo-*** means 1,000 times as large, so 1,000 seconds = 1 **kilo**second.
b. The prefix ***micro-*** means one-millionth as large (0.000 001), so 0.000 001 gram = 1 **micro**gram.

Adam Gault /Science Photo Library RF/Science Source

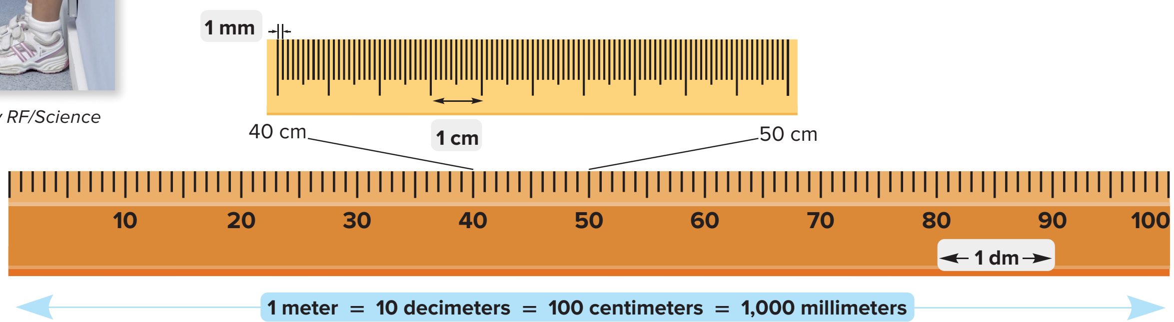
Practice Problem 1.2

What term is used for each of the following units: (a) a million liters; (b) a thousandth of a second; (c) a hundredth of a gram; (d) a tenth of a liter?

More Practice: Try Problem 1.8.

1.4B Measuring Length

The base unit of length in the metric system is the *meter* (m). A meter, 39.4 inches in the English system, is slightly longer than a yard (36 inches). Common units derived from a meter, shown with a size comparison on a meter stick, are the decimeter (dm), centimeter (cm), and millimeter (mm).



Note how these values are related to those in Table 1.2. Because a centimeter is one *hundredth* of a meter (0.01 m), there are *100* centimeters in a meter.

Problem 1.7

If a nanometer is one billionth of a meter (0.000 000 001 m), how many nanometers are there in one meter?

Photodisc/Getty Images

1.4C Measuring Mass

Although the terms mass and weight are often used interchangeably, they really have different meanings.

- *Mass* is a measure of the amount of matter in an object.
- *Weight* is the force that matter feels due to gravity.

The mass of an object is independent of its location. The weight of an object changes slightly with its location on the earth, and drastically when the object is moved from the earth to the moon, where the gravitational pull is only one-sixth that of the earth. Although we often speak of *weighing* an object, we are really *measuring its mass.*

The base unit of mass in the metric system is the *gram* (g), a small quantity compared to the English pound (1 lb = 454 g). Two common units derived from a gram are the kilogram (kg) and milligram (mg).

$$1\ \text{kg} = 1{,}000\ \text{g}$$
$$1\ \text{g} = 1{,}000\ \text{mg}$$

Problem 1.8

What term is used for each quantity? (For example, 0.001 g = one milligram.)

a. 0.000 001 g
b. 1,000,000,000 m
c. 0.000 000 001 s
d. 0.01 g

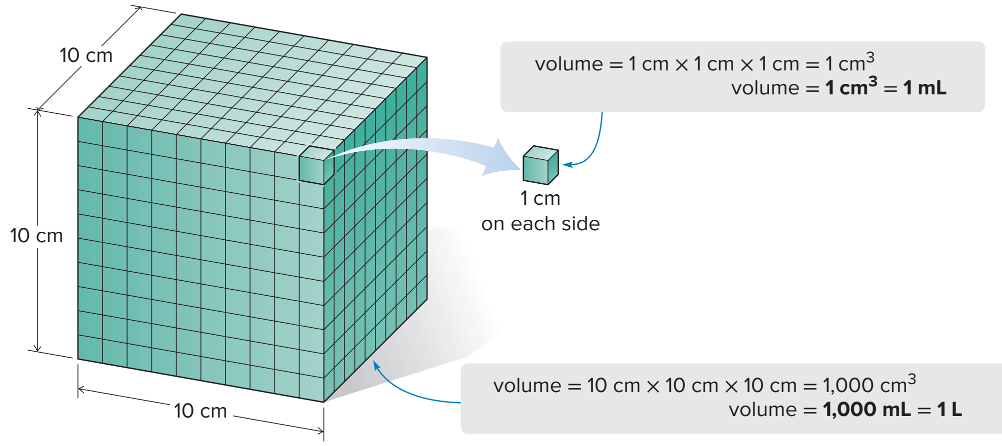
Note the difference between the units **cm** and **cm³.** The centimeter (cm) is a unit of *length.* A cubic centimeter (cm^3 or cc) is a unit of *volume.*

Rawpixel/123RF

1 mL = 1 cm^3 = 1 cc

1.4D Measuring Volume

The base unit of volume in the metric system is the *liter* (L), which is slightly larger than the English quart (1 L = 1.06 qt). One liter is defined as the volume of a cube 10 cm on an edge.



Three common units derived from a liter used in medicine and laboratory research are the deciliter (dL), milliliter (mL), and microliter (μL). **One milliliter is the same as one cubic centimeter (cm^3), which is abbreviated as cc.**

1 L = 10 dL
1 L = 1,000 mL
1 L = 1,000,000 μL

Table 1.3 summarizes common metric units of length, mass, and volume. Table 1.4 lists English units of measurement, as well as their metric equivalents.

Problem 1.9

Using the prefixes in Table 1.2, determine which quantity in each pair is larger.

a. 3 mL or 3 cL
b. 1 ng or 1 μg
c. 5 km or 5 cm
d. 2 mL or 2 μL

Table 1.3 Summary of the Common Metric Units of Length, Mass, and Volume

Length	Mass	Volume
1 km = 1,000 m	1 kg = 1,000 g	1 L = 10 dL
1 m = 100 cm	1 g = 1,000 mg	1 L = 1,000 mL
1 m = 1,000 mm	1 mg = 1,000 μg	1 L = 1,000,000 μL
1 cm = 10 mm		1 dL = 100 mL
		1 mL = 1 cm^3 = 1 cc

Common metric abbreviations and conversion factors are also listed in Appendix C.

Table 1.4 English Units and Their Metric Equivalents

Quantity	English Unit	Metric–English Relationship
Length	1 ft = 12 in.	2.54 cm = 1 in.
	1 yd = 3 ft	1 m = 39.4 in.
	1 mi = 5,280 ft	1 km = 0.621 mi
Mass	1 lb = 16 oz	1 kg = 2.20 lb
	1 ton = 2,000 lb	454 g = 1 lb
		28.3 g = 1 oz
Volume	1 qt = 4 cups	946 mL = 1 qt
	1 qt = 2 pt	1 L = 1.06 qt
	1 qt = 32 fl oz	29.6 mL = 1 fl oz
	1 gal = 4 qt	

Common abbreviations for English units: inch (in.), foot (ft), yard (yd), mile (mi), pound (lb), ounce (oz), gallon (gal), quart (qt), pint (pt), and fluid ounce (fl oz).

A container of 71 macadamia nuts weighs 125 g. The number of nuts (71) is exact, whereas the mass of the nuts (125 g) is inexact. *Zachary D.-K. Smith*

1.5 Significant Figures

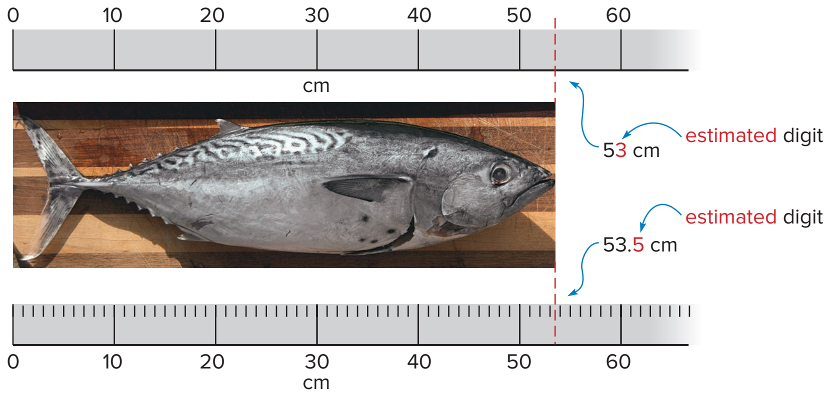
Numbers used in chemistry are either **exact** or **inexact.**

- **An *exact* number results from counting objects or is part of a definition.**

Our bodies have 10 fingers, 10 toes, and two kidneys. A meter is composed of 100 centimeters. These numbers are exact because there is no uncertainty associated with them.

- **An *inexact* number results from a measurement or observation and contains some uncertainty.**

Whenever we measure a quantity there is a degree of uncertainty associated with the result. The last number (farthest to the right) is an estimate, and it depends on the type of measuring device we use to obtain it. For example, the length of a fish caught on a recent outing could be reported as 53 cm or 53.5 cm depending on the tape measure used.



Daniel C. Smith

- ***Significant figures* are all the digits in a measured number including *one* estimated digit.**

Thus, the length 53 cm has two significant figures, and the length 53.5 cm has three significant figures.

1.5A Determining the Number of Significant Figures

How many significant figures are contained in a number?

- **All nonzero digits are always significant.**

65.2 g	**three** significant figures
1,265 m	**four** significant figures
25 µL	**two** significant figures
255.345 g	**six** significant figures

Whether a zero counts as a significant figure depends on its location in the number.

Rules to Determine When a Zero Is a Significant Figure

Rule [1] A zero *counts* as a significant figure when it occurs:

• Between two nonzero digits	29.05 g—four significant figures 1.0087 mL—five significant figures
• At the end of a number with a decimal point	25.70 cm—four significant figures 3.7500 g—five significant figures 620. lb—three significant figures

Rule [2] A zero does *not* count as a significant figure when it occurs:

• At the beginning of a number	0.0245 mg—three significant figures 0.008 mL—one significant figure
• At the end of a number that does *not* have a decimal point	2,570 m—three significant figures 1,245,500 m—five significant figures

In reading a number with a decimal point from left to right, **all digits starting with the first nonzero number are significant figures.** The number 0.003 450 120 has seven significant figures, shown in red.

Sample Problem 1.3 Determining the Number of Significant Figures

How many significant figures does each number contain?

a. 34.08 b. 0.0054 c. 260.00 d. 260

Analysis

All nonzero digits are significant. A zero is significant when it occurs between two nonzero digits, or at the end of a number with a decimal point.

Solution

Significant figures are shown in red.

a. 34.08 (four) b. 0.0054 (two) c. 260.00 (five) d. 260 (two)

Practice Problem 1.3

How many significant figures does each number contain?

a. 23.45 c. 230 e. 0.202 g. 1,245,006 i. 10,040

b. 23.057 d. 231.0 f. 0.003 60 h. 1,200,000 j. 10,040.

More Practice: Try Problems 1.33, 1.34.

Problem 1.10

Indicate whether each zero in the following numbers is significant.

a. 0.003 04 b. 26,045 c. 1,000,034 d. 0.304 00

1.5B Using Significant Figures in Multiplication and Division

We often must perform calculations with numbers that contain a different number of significant figures. The number of significant figures in the answer of a problem depends on the type of mathematical calculation—multiplication (and division) or addition (and subtraction).

- **In multiplication and division, the answer has the same number of significant figures as the original number with the *fewest* significant figures.**

Let's say you drove a car 351.2 miles in 5.5 hours, and you wanted to calculate how many miles per hour you traveled. Entering these numbers on a calculator would give the following result:

four significant figures

$$\text{Miles per hour} = \frac{351.2 \text{ miles}}{5.5 \text{ hours}} = 63.854\,545 \text{ miles per hour}$$

two significant figures

The answer must contain only two significant figures.

The answer to this problem can have only *two* significant figures, because one of the original numbers (5.5 hours) has only *two* significant figures. To write the answer in proper form, we must **round off the number** to give an answer with only two significant figures. Two rules are used in rounding off numbers.

- **If the first number that must be dropped is 4 or fewer, *drop it and all remaining numbers.***
- **If the first number that must be dropped is 5 or greater, *round the number up* by adding one to the last digit that will be retained.**

In this problem:

These digits must be retained.

first digit to be dropped

63.854 545

These digits must be dropped.

- Because the first digit to be dropped is **8** (5 or greater), add 1 to the digit to its left.
- The answer 63.854 545 rounded to **two** digits is **64 miles per hour.**

Table 1.5 gives other examples of rounding off numbers.

Table 1.5 Rounding Off Numbers

Original Number	Rounded To	Rounded Number
61.2537	Two places	61
61.2537	Three places	61.3
61.2537	Four places	61.25
61.2537	Five places	61.254

The first number to be dropped is indicated in red in each original number. When this number is 4 or fewer, it and all other digits to its right are dropped. When this number is 5 or greater, 1 is added to the digit to its left.

A scientific calculator is useful in simple mathematical calculations, as well as more complicated functions, such as converting a number to scientific notation (Section 1.6B) or taking the logarithm of a number (Section 9.6B). *S.Narongrit/Shutterstock*

When a calculator is used in calculations, sometimes the display shows *more* digits than are significant and sometimes it shows *fewer* digits. For example, in multiplying 23.2 by 1.1, the calculator displays the answer as 25.52. Because the quantity 1.1 has only two significant figures, the answer must contain only two significant figures and be rounded to 26.

23.2	×	1.1	=	25.52	round off →	26
three significant figures		**two** significant figures		**calculator display**		**two** significant figures
						Answer

In contrast, dividing 25.0 by 0.50 displays the answer as 50, a quantity with only one significant figure. Because 0.50 has *two* significant figures, the answer must contain *two* significant figures, and this is achieved by adding a decimal point (50.).

25.0	×	0.50	=	50	add decimal →	50.
three significant figures		**two** significant figures		**calculator display**		**two** significant figures
						Answer

Sample Problem 1.4 Rounding Off Numbers

Round off each number to three significant figures.

a. 1.2735 b. 0.002 536 22 c. 3,836.9

Analysis

If the answer is to have *three* significant figures, look at the *fourth* number from the left.

- If this number is 4 or fewer, *drop it* and all remaining numbers to the right.
- If the fourth number from the left is 5 or greater, *round the number up* by adding one to the third digit.

Solution

a. 1.27 b. 0.002 54 c. 3,840 (Omit the decimal point after the 0. The number 3,840. has four significant figures.)

Practice Problem 1.4

Round off each number in Sample Problem 1.4 to two significant figures.

More Practice: Try Problems 1.35, 1.36.

Sample Problem 1.5 Determining Significant Figures in Multiplication and Division

Carry out each calculation and give the answer using the proper number of significant figures.

a. 3.81×0.046 b. $120.085 \div 106$

Analysis

Because these calculations involve multiplication and division, the answer must have the same number of significant figures as the original number with the *fewer* number of significant figures.

Solution

a. $3.81 \times 0.046 = 0.1753$

- Because 0.046 has only two significant figures, round the answer to give it two significant figures.

0.1753 Because this number is 5 (5 or greater), round the 7 to its left up by one.

Answer: 0.18

b. 120.085 ÷ 106 = 1.132 877 36

- Because 106 has three significant figures, round the answer to give it three significant figures.

1.132 877 36 Because this number is 2 (4 or fewer), drop it and all numbers to its right.

Answer: 1.13

Practice Problem 1.5

Carry out each calculation and give the answer using the proper number of significant figures.

a. 10.70 × 3.5 b. 0.206 ÷ 25,993 c. 1,300 ÷ 41.2 d. 120.5 × 26

More Practice: Try Problems 1.37a, c, e; 1.38a, c, e.

1.5C Using Significant Figures in Addition and Subtraction

In determining significant figures in addition and subtraction, the decimal place of the last significant digit determines the number of significant figures in the answer.

- **In addition and subtraction, the answer has the same number of decimal places as the original number with the *fewest* decimal places.**

Suppose a baby weighed 3.6 kg at birth and 10.11 kg on his first birthday. To figure out how much weight the baby gained in his first year of life, we subtract these two numbers and report the answer using the proper number of significant figures.

weight at one year = 10.11 kg — 10.11 kg **two** digits after the decimal point

weight at birth = 3.6 kg — – 3.6 kg **one** digit after the decimal point

weight gain = 6.51 kg

last significant digit

- The answer can have only **one** digit after the decimal point.
- Round 6.51 to 6.5.
- The baby gained 6.5 kg during his first year of life.

Because 3.6 kg has only one significant figure after the decimal point, the answer can have only one significant figure after the decimal point as well.

Sample Problem 1.6 Determining Significant Figures in Addition and Subtraction

While on a diet, a woman lost 3.52 lb the first week, 2.2 lb the second week, and 0.59 lb the third week. How much weight did she lose in all?

Analysis

Add up the amount of weight loss each week to get the total weight loss. When adding, the answer has the same number of decimal places as the original number with the fewest decimal places.

Solution

3.52 lb

2.2 lb one digit after the decimal point

0.59 lb

6.31 lb ------→ 6.3 lb
round off

last significant digit

- Because 2.2 lb has only one digit after the decimal point, the answer can have only one digit after the decimal point.
- Round 6.31 to 6.3.
- Total weight loss: 6.3 lb.

Practice Problem 1.6

Carry out each calculation and give the answer using the proper number of significant figures.

a. 27.8 cm + 0.246 cm
b. 102.66 mL + 0.857 mL + 24.0 mL
c. 54.6 mg – 25 mg
d. 2.35 s – 0.266 s

More Practice: Try Problems 1.37b, d, f; 1.38b, d, f.

1.6 Scientific Notation

Healthcare professionals and scientists must often deal with very large and very small numbers. For example, the blood platelet count of a healthy adult might be 250,000 platelets per mL. At the other extreme, the level of the female sex hormone estriol during pregnancy might be 0.000 000 250 g per mL of blood plasma. Estriol is secreted by the placenta and its concentration is used as a measure of the health of the fetus.

1.6A Writing Numbers in Scientific Notation

To write numbers that contain many leading zeros (at the beginning) or trailing zeros (at the end), scientists use **scientific notation.**

- **In scientific notation, a number is written as $y \times 10^x$.**
- **The term y, called the coefficient, is a number between 1 and 10.**
- **The value x is an exponent, which can be any positive or negative whole number.**

Hospital laboratory technicians determine thousands of laboratory results each day. *Stockbyte/Getty Images*

First, let's recall what powers of 10 with *positive* exponents, such as 10^2 or 10^5, mean. These correspond to numbers greater than one, and the positive exponent tells how many zeros are to be written after the number one. Thus, $10^2 = 100$, a number with two zeros after the number one.

The product has **two** zeros.

$$10^{2} = 10 \times 10 = 100$$

The exponent **2** means "multiply **two** 10s."

The product has **five** zeros.

$$10^{5} = 10 \times 10 \times 10 \times 10 \times 10 = 100{,}000$$

The exponent **5** means "multiply **five** 10s."

Powers of 10 that contain *negative* exponents, such as 10^{-3}, correspond to numbers less than one. In this case the exponent tells how many places (*not* zeros) are located to the right of the decimal point.

The answer has **three** places to the *right* of the decimal point, including the number one.

$$10^{-3} = \frac{1}{10 \times 10 \times 10} = 0.001$$

The exponent **–3** means "divide by **three** 10s."

To write a number in scientific notation, we follow a stepwise procedure.

How To Convert a Standard Number to Scientific Notation

Example: Write each number in scientific notation: (a) 2,500; (b) 0.036.

Step [1] **Move the decimal point to give a number between 1 and 10.**

a. 2500.

Move the decimal point three places to the left to give the number 2.5.

b. 0.036

Move the decimal point two places to the right to give the number 3.6.

—Continued

***How To*, continued . . .**

Step [2] **Multiply the result by 10^x, where *x* is the number of places the decimal point was moved.**

- If the decimal point is moved to the **left,** *x* is **positive.**
- If the decimal point is moved to the **right,** *x* is **negative.**

a. Since the decimal point was moved three places to the **left,** the exponent is +3, and the coefficient is multiplied by 10^3.

Answer: $2{,}500 = 2.5 \times 10^3$

b. Since the decimal point was moved two places to the **right,** the exponent is −2, and the coefficient is multiplied by 10^{-2}.

Answer: $0.036 = 3.6 \times 10^{-2}$

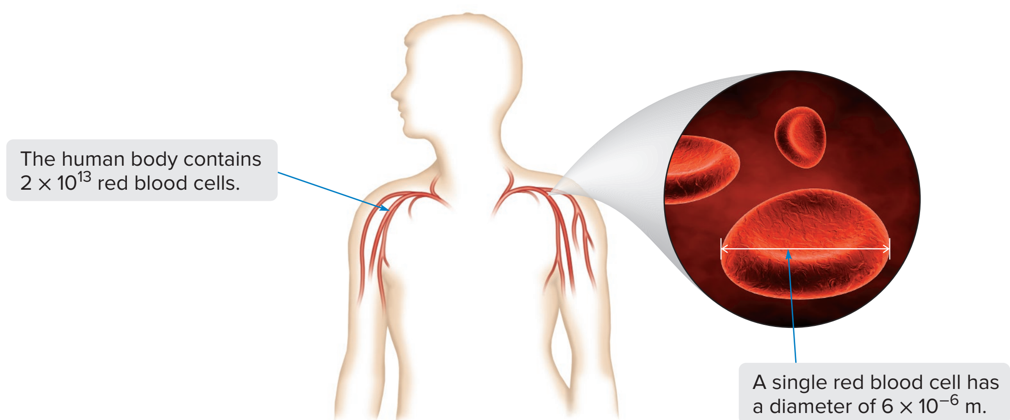
Notice that the number of significant figures in the coefficient in scientific notation must equal the number of significant figures in the original number. Thus, the coefficients for both 2,500 and 0.036 need two significant figures and no more. Figure 1.7 shows two more examples of numbers written in standard form and scientific notation.

$2{,}500 = 2.5 \times 10^3$ — two significant figures

not 2.50×10^3 (three significant figures)

not 2.500×10^3 (four significant figures)

Figure 1.7 Numbers in Standard Form and Scientific Notation



Cre8tive Studios/Alamy Stock Photos

Very large and very small numbers are more conveniently written in scientific notation.

Quantity	Number	Scientific Notation
Number of red blood cells	20,000,000,000,000	2×10^{13}
Diameter of a red blood cell	0.000 006 m	6×10^{-6} m

Sample Problem 1.7 Writing a Number in Scientific Notation

Write the recommended daily dietary intake of each nutrient in scientific notation: (a) sodium, 2,400 mg; (b) vitamin B_{12}, 0.000 006 g.

Analysis

Move the decimal point to give a number between 1 and 10. Multiply the number by 10^x, where *x* is the number of places the decimal point was moved. The exponent *x* is **(+)** when the decimal point moves to the ***left*** and **(–)** when it moves to the ***right.***

Solution

a.

$2400. = 2.4 \times 10^3$ the number of places the decimal point was moved to the left

Move the decimal point three places to the left.

- Write the coefficient as 2.4 (two significant figures), because 2,400 contains two significant figures.

b.

$0.000\ 006 = 6 \times 10^{-6}$ the number of places the decimal point was moved to the right

Move the decimal point six places to the right.

- Write the coefficient as 6 (one significant figure), because 0.000 006 contains one significant figure.

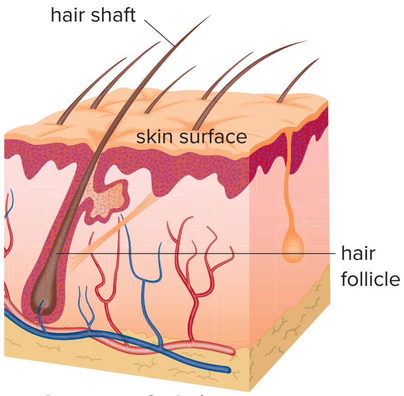
Practice Problem 1.7

Write each number in scientific notation.

a. 93,200 c. 6,780,000 e. 4,520,000,000,000

b. 0.000 725 d. 0.000 030 f. 0.000 000 000 028

More Practice: Try Problems 1.39, 1.40, 1.45.



Anatomy of a hair

Hair is a protein that grows from follicles in the skin. We will learn more about the proteins in hair in Section 21.7.

To convert a number in scientific notation to a standard number, reverse the procedure, as shown in Sample Problem 1.8. It is often necessary to add leading or trailing zeros to write the number.

- **When the exponent *x* is positive, move the decimal point *x* places to the *right*.**

2.800×10^{2} 2.800 ---→ 280.0

Move the decimal point to the *right* **two** places.

- **When the exponent *x* is negative, move the decimal point *x* places to the *left*.**

2.80×10^{-2} 002.80 ---→ 0.0280

Move the decimal point to the *left* **two** places.

Sample Problem 1.8 Converting a Number in Scientific Notation to a Standard Number

The average diameter of a human hair is about 1.2×10^{-4} m. Convert this value to a standard number.

Analysis

The exponent in 10^x tells how many places to move the decimal point in the coefficient to generate a standard number. The decimal point goes to the *right* when *x* is positive and to the *left* when *x* is negative.

Solution

1.2×10^{-4} 000 01.2 ---→ 0.000 12 m

Move the decimal point to the *left* four places. **Answer**

The answer, 0.000 12, has two significant figures, just like 1.2×10^{-4}.

Practice Problem 1.8

Convert each number to its standard form.

a. 6.5×10^3 c. 3.780×10^{-2} e. 2.221×10^6

b. 3.26×10^{-5} d. 1.04×10^8 f. 4.5×10^{-10}

More Practice: Try Problems 1.41, 1.42.

1.6B Using a Scientific Calculator for Numbers in Scientific Notation

Several examples involving the use of a scientific calculator are found in Appendix D, Useful Mathematical Concepts.

A scientific calculator can be used to convert a standard number to a number in scientific notation. Described in this section are the steps that can be followed in calculations with *some* calculators. Consult your manual if these steps do *not* produce the stated result.

To convert a number from its standard form to a number in scientific notation, follow the steps in columns [1]–[3].

	1 Number to enter	2 Buttons to press	3 Display	Meaning
A number ≥ 1:	1,200	**2nd** **SCI**	1.2^{03}	1.2×10^3
A **decimal** < 1:	0.052	**2nd** **SCI**	5.2^{-02}	5.2×10^{-2}

To enter a number in scientific notation, follow the steps in columns [1]–[3].

	1 Number to enter	2 Buttons to press	3 Display	Meaning
A number with a **positive** exponent, as in 1.5×10^8:	1.5	**EE** **8**	1.5^{08}	150,000,000
A number with a **negative** exponent, as in 3.5×10^{-4}:	3.5	**EE** **4** **Change sign (+ to –)**	3.5^{-04}	0.000 35

Problem 1.11

If you used the calculator described in this section, write each of the following calculator displays as a number in standard form and using scientific notation: (a) 5.43^{05}; (b) 4.43^{-06}; (c) 7.4^{04}.

1.7 Problem Solving Using Conversion Factors

Often a measurement is recorded in one unit, and then it must be converted to another unit. For example, a patient may weigh 130 lb, but we may need to know her weight in kilograms to calculate a drug dosage. The recommended daily dietary intake of potassium is 3,500 mg, but we may need to know how many grams this corresponds to.

1.7A Conversion Factors

To convert one unit to another we use one or more **conversion factors.**

original quantity × conversion factor = desired quantity

These two quantities are equivalent.
Only the units are different.

- A *conversion factor* is a term that converts a quantity in one unit to a quantity in another unit.

A conversion factor is formed by taking an equality, such as 2.20 lb = 1 kg, and writing it as a ratio. We can always write a conversion factor in two different ways.

$$\frac{2.20\ \text{lb}}{1\ \text{kg}} \quad \text{or} \quad \frac{1\ \text{kg}}{2.20\ \text{lb}}$$

numerator

denominator

conversion factors for pounds and kilograms

Refer to Tables 1.3 and 1.4 for metric and English units needed in problem solving. Common metric and English units are also listed in Appendix C.

With pounds and kilograms, either of these values can be written above the division line of the fraction (the **numerator**) or below the division line (the **denominator**). The way the conversion factor is written will depend on the problem.

Sample Problem 1.9 Writing Conversion Factors for a Pair of Units

Write two conversion factors for each pair of units: (a) kilograms and grams; (b) quarts and liters.

Analysis

Use the equalities in Tables 1.3 and 1.4 to write a ratio that shows the relationship between the two units.

Solution

a. Conversion factors for kilograms and grams:

$$\frac{1000\text{ g}}{1\text{ kg}} \quad \text{or} \quad \frac{1\text{ kg}}{1000\text{ g}}$$

b. Conversion factors for quarts and liters:

$$\frac{1.06\text{ qt}}{1\text{ L}} \quad \text{or} \quad \frac{1\text{ L}}{1.06\text{ qt}}$$

Practice Problem 1.9

Write two conversion factors for each pair of units.

a. miles and kilometers
b. meters and millimeters
c. grams and pounds
d. milligrams and micrograms

1.7B Solving a Problem Using One Conversion Factor

When using conversion factors to solve a problem, if a unit appears in the numerator in one term and the denominator in another term, the units *cancel*. **The goal in setting up a problem is to make sure *all unwanted units cancel.***

Let's say we want to convert 130 lb to kilograms.

$$\underset{\text{original quantity}}{130\text{ lb}} \times \boxed{\text{conversion factor}} = \underset{\text{desired quantity}}{?\ \text{kg}}$$

Two possible conversion factors: $\dfrac{2.20\text{ lb}}{1\text{ kg}}$ or $\dfrac{1\text{ kg}}{2.20\text{ lb}}$

To solve this problem we must use a conversion factor that satisfies two criteria.

- **The conversion factor must contain the two quantities in question—pounds and kilograms.**
- **The conversion factor must cancel out the unwanted unit—pounds.**

This means choosing the conversion factor with the unwanted unit—pounds—*in the denominator* to cancel out pounds in the original quantity. This leaves kilograms as the only remaining unit, and the problem is solved.

conversion factor

$$130\ \cancel{\text{lb}} \times \frac{1\text{ kg}}{2.20\ \cancel{\text{lb}}} = \boxed{59\text{ kg}} \quad \textbf{answer in kilograms}$$

Pounds (lb) must be the denominator to cancel the unwanted unit (lb) in the original quantity.

How many grams of aspirin are contained in a 325-mg tablet? *Mark Dierker/McGraw-Hill*

We must use the correct number of significant figures in reporting an answer to each problem. In this case, the value 1 kg is *defined* as 2.20 lb; in other words, 1 kg contains the exact number "1" with *no* uncertainty, so it does not limit the number of digits in the answer. Since 130 lb has two significant figures, the answer is rounded to two significant figures (59 kg).

As problems with units get more complicated, keep in mind the following general steps that are useful for solving any problem using conversion factors.

How To Solve a Problem Using Conversion Factors

Example: How many grams of aspirin are contained in a 325-mg tablet?

Step [1] **Identify the original quantity and the desired quantity, including units.**

- In this problem the original quantity is reported in milligrams and the desired quantity is in grams.

325 mg	? g
original quantity	desired quantity

Step [2] **Write out the conversion factor(s) needed to solve the problem.**

- We need a conversion factor that relates milligrams and grams (Table 1.3). Since the unwanted unit is in milligrams, **choose the conversion factor that contains milligrams in the denominator so that the *units cancel.***

Two possible conversion factors: $\dfrac{1000 \text{ mg}}{1 \text{ g}}$ or $\boxed{\dfrac{1 \text{ g}}{1000 \text{ mg}}}$ Choose this factor to cancel the unwanted unit, mg.

- Sometimes one conversion factor is all that is needed in a problem. At other times (Section 1.7C) more than one conversion factor is needed.
- If the desired answer has a single unit (grams in this case), **the conversion factor must contain the desired unit in the numerator and the unwanted unit in the denominator.**

Step [3] **Set up and solve the problem.**

- Multiply the original quantity by the conversion factor to obtain the desired quantity.

conversion factor

$$\underset{\text{original quantity}}{325 \cancel{\text{mg}}} \times \frac{1 \text{ g}}{1000 \cancel{\text{mg}}} = \underset{\text{desired quantity}}{0.325 \text{ g of aspirin}}$$

The number of mg (unwanted unit) cancels.

Step [4] **Write the answer using the correct number of significant figures and check it by estimation.**

- Use the number of significant figures in each inexact (measured) number to determine the number of significant figures in the answer. In this case the answer is limited to three significant figures by the original quantity (325 mg).
- Estimate the answer using a variety of methods. In this case we knew our answer had to be less than one, because it is obtained by dividing 325 by a number larger than itself.

Sample Problem 1.10 Solving a Problem Using Conversion Factors

If 1.65×10^9 lb of PET (polyethylene terephthalate), the plastic used in water and soft drink bottles, is recycled in a year, how many kilograms of PET does this represent?

ENVIRONMENTAL NOTE

As we will learn in Chapter 13, PET, the plastic used in water bottles, can be recycled into bags, containers, fleece jackets, and carpeting. *Angela Hampton Picture Library/Alamy Stock Photo*

Analysis and Solution

[1] Identify the original quantity and the desired quantity.

1.65×10^9 lb	? kg
original quantity	desired quantity

[2] Write out the conversion factors.

- We need a conversion factor that relates pounds and kilograms (Table 1.4). Choose the conversion factor that places pounds in the denominator, so that the units cancel.

$\dfrac{2.20 \text{ lb}}{1 \text{ kg}}$ or $\boxed{\dfrac{1 \text{ kg}}{2.20 \text{ lb}}}$ Choose this factor to cancel the unwanted unit, lb.

STUDY TIP

Whenever a calculation requires conversion of a quantity from one unit to another, follow the three-step *How To:*

1. Identify the units in the original quantity and the desired quantity.
2. Write out the conversion factor(s).
3. Arrange the conversion factors so that unwanted units cancel, then solve.

[3] Solve the problem.

- Multiply the original quantity by the conversion factor to obtain the desired quantity. Set up the problem, so that the unwanted unit, pounds, cancels.

conversion factor

$$1.65 \times 10^9\ \cancel{\text{lb}} \times \frac{1\ \text{kg}}{2.20\ \cancel{\text{lb}}} = 7.50 \times 10^8\ \text{kg of PET}$$

The number of lb (unwanted unit) cancels.

Practice Problem 1.10

Carry out each of the following conversions.

a. 25 L to dL   b. 40.0 oz to g   c. 32 in. to cm   d. 10 cm to mm

More Practice: Try Problems 1.47–1.49, 1.52, 1.53a, 1.54a.

Problem 1.12

If the mass of the baby in the chapter-opening photograph is 9 lb 8.0 oz, how many kilograms does the baby weigh?

Problem 1.13

(a) What is the volume of liquid contained in the given 0.5-mL syringe? (b) Convert this value to microliters.

1.7C Solving a Problem Using Two or More Conversion Factors

Some problems require the use of more than one conversion factor to obtain the desired units in the answer. The same stepwise procedure is followed no matter how many conversion factors are needed. Keep in mind:

- **Always arrange the factors so that the denominator in one term *cancels* the numerator in the preceding term.**

Sample Problem 1.11 illustrates how to solve a problem with two conversion factors.

Sample Problem 1.11 **Solving a Problem Using More Than One Conversion Factor**

An individual donated 1.0 pt of blood at the local blood bank. How many liters of blood does this correspond to?

How many liters of blood does this pint of blood contain? *Keith Brofsky/Photodisc/Getty Images*

Analysis and Solution

[1] Identify the original quantity and the desired quantity.

1.0 pt	? L
original quantity	desired quantity

[2] Write out the conversion factors.

- We have no conversion factor that relates pints to liters directly. We do, however, know conversions for pints to quarts, and quarts to liters.

pint–quart conversion

$$\frac{2\ \text{pt}}{1\ \text{qt}} \quad \text{or} \quad \boxed{\frac{1\ \text{qt}}{2\ \text{pt}}}$$

quart–liter conversion

$$\frac{1.06\ \text{qt}}{1\ \text{L}} \quad \text{or} \quad \boxed{\frac{1\ \text{L}}{1.06\ \text{qt}}}$$

Choose the conversion factors with the unwanted units—pt and qt—in the denominator.

[3] Solve the problem.

- To set up the problem so that unwanted units cancel, arrange each term so that the **units in the numerator of one term *cancel* the units in the denominator of the adjacent term.** In this problem we need to cancel both pints and quarts to get liters.
- The single desired unit, liters, must be in the **numerator** of one term.

Liters do not cancel.

$$1.0\ \cancel{\text{pt}} \times \frac{1\ \cancel{\text{qt}}}{2\ \cancel{\text{pt}}} \times \frac{1\ \text{L}}{1.06\ \cancel{\text{qt}}} = 0.47\ \text{L}$$

Pints cancel. Quarts cancel.

[4] Check.

- Because there are two pints in a quart and a quart is about the same size as a liter, one pint should be about half a liter. The answer, 0.47, is just about 0.5.
- Write the answer with two significant figures because one term, 1.0 pt, has two significant figures.

Practice Problem 1.11

Carry out each of the following conversions.

a. 6,250 ft to km b. 3 cups to L c. 4.5 ft to cm

More Practice: Try Problems 1.51, 1.53b, 1.54b, 1.70, 1.72–1.74, 1.76, 1.77.

1.8 FOCUS ON HEALTH & MEDICINE
Problem Solving Using Clinical Conversion Factors

HEALTH NOTE

The active ingredient in Children's Tylenol is acetaminophen. *Jill Braaten/McGraw-Hill*

Sometimes conversion factors don't have to be looked up in a table; they are stated in the problem. If a drug is sold as a 250-mg tablet, this fact becomes a conversion factor relating milligrams to tablets.

$$\frac{250\ \text{mg}}{1\ \text{tablet}} \quad \text{or} \quad \frac{1\ \text{tablet}}{250\ \text{mg}}$$

mg–tablet conversion factors

Alternatively, a drug could be sold as a liquid solution with a specific concentration. For example, Children's Tylenol contains 80 mg of the active ingredient acetaminophen in 2.5 mL. This fact becomes a conversion factor relating milligrams to milliliters.

$$\frac{80\ \text{mg}}{2.5\ \text{mL}} \quad \text{or} \quad \frac{2.5\ \text{mL}}{80\ \text{mg}}$$

mg of acetaminophen–mL conversion factors

Sample Problems 1.12 and 1.13 illustrate how these conversion factors are used in determining drug dosages.

Sample Problem 1.12 — **Using Clinical Conversion Factors to Solve a Problem**

A patient is prescribed 1.25 g of amoxicillin, which is available in 250-mg tablets. How many tablets are needed?

Analysis and Solution

[1] Identify the original quantity and the desired quantity.

- We must convert the number of grams of amoxicillin needed to the number of tablets that must be administered.

1.25 g	? tablets
original quantity	desired quantity

[2] Write out the conversion factors.

- We have no conversion factor that relates grams to tablets directly. We do know, however, how to relate grams to milligrams, and milligrams to tablets.

g–mg conversion factors

$$\frac{1\text{ g}}{1000\text{ mg}} \quad \text{or} \quad \boxed{\frac{1000\text{ mg}}{1\text{ g}}}$$

mg–tablet conversion factors

$$\frac{250\text{ mg}}{1\text{ tablet}} \quad \text{or} \quad \boxed{\frac{1\text{ tablet}}{250\text{ mg}}}$$

Choose the conversion factors with the unwanted units–g and mg–in the denominator.

[3] Solve the problem.

- Arrange each term so that the **units in the numerator of one term *cancel* the units in the denominator of the adjacent term.** In this problem we need to cancel both grams and milligrams to get tablets.
- The single desired unit, tablets, must be located in the **numerator** of one term.

Tablets do not cancel.

$$1.25\ \cancel{\text{g}} \times \frac{1000\ \cancel{\text{mg}}}{1\ \cancel{\text{g}}} \times \frac{1\text{ tablet}}{250\ \cancel{\text{mg}}} = 5\text{ tablets}$$

Grams cancel. Milligrams cancel.

[4] Check.

- The answer of 5 tablets of amoxicillin (not 0.5 or 50) is reasonable. Since the dose in a single tablet (250 mg) is a fraction of a gram, and the required dose is more than a gram, the answer must be greater than one.

Practice Problem 1.12

A patient is prescribed 0.100 mg of a drug that is available in 25-µg tablets. How many tablets are needed?

More Practice: Try Problem 1.79.

Sample Problem 1.13 Using a Clinical Conversion Factor to Solve a Problem

A dose of 240 mg of acetaminophen is prescribed for a 20-kg child. How many mL of Children's Tylenol (80. mg of acetaminophen per 2.5 mL) are needed?

Analysis and Solution

[1] Identify the original quantity and the desired quantity.

- We must convert the number of milligrams of acetaminophen needed to the number of mL that must be administered.

240 mg	? mL
original quantity	desired quantity

[2] Write out the conversion factors.

mg of acetaminophen–mL conversion factors

$$\frac{80.\text{ mg}}{2.5\text{ mL}} \quad \text{or} \quad \boxed{\frac{2.5\text{ mL}}{80.\text{ mg}}}$$

Choose the conversion factor to cancel mg.

[3] Solve the problem.

- Arrange the terms so that the **units in the numerator of one term *cancel* the units of the denominator of the adjacent term.** In this problem we need to cancel milligrams to obtain milliliters.

- In this problem we are given a fact we don't need to use—the child weighs 20 kg. We can ignore this quantity in carrying out the calculation.

$$240\ \cancel{\text{mg}} \times \frac{2.5\ \text{mL}}{80.\ \cancel{\text{mg}}} = 7.5\ \text{mL of Children's Tylenol}$$

Milligrams cancel.

[4] Check.

- The answer of 7.5 mL (not 0.75 or 75) is reasonable. Because the required dose is larger than the dose in 2.5 mL, the answer must be larger than 2.5 mL.

Mark Dierker/McGraw-Hill

Practice Problem 1.13

(a) How many milliliters are contained in the dose of Children's Tylenol shown in the adjacent photo (1 teaspoon = 5 mL)? (b) If Children's Tylenol contains 80. mg of acetaminophen per 2.5 mL, how much acetaminophen (in mg) is contained in the dose?

More Practice: Try Problems 1.50, 1.69, 1.71, 1.75, 1.78, 1.80.

Problem 1.14

How many milliliters of Children's Motrin (100 mg of ibuprofen per 5 mL) are needed to give a child a dose of 160 mg?

1.9 Temperature

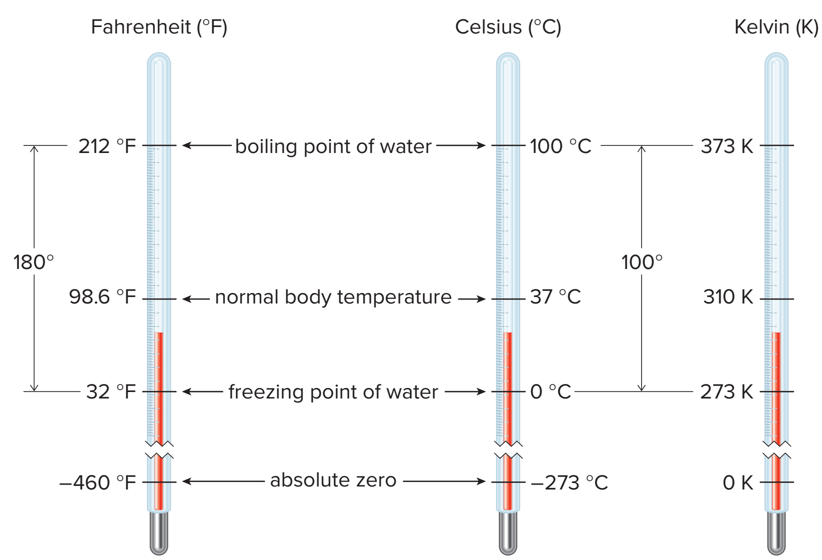
***Temperature* is a measure of how hot or cold an object is.** Three temperature scales are used: **Fahrenheit** (most common in the United States), **Celsius** (most commonly used by scientists and countries other than the United States), and **Kelvin** (Figure 1.8).

The Fahrenheit and Celsius scales are both divided into **degrees.** On the Fahrenheit scale, water freezes at 32 °F and boils at 212 °F. On the Celsius scale, water freezes at 0 °C and boils at

CONSUMER NOTE

Although mercury thermometers were used in hospitals to measure temperature for many years, temperature is now more commonly recorded with a digital thermometer. Tympanic thermometers, which use an infrared sensing device placed in the ear, are also routinely used.
Jill Braaten/McGraw-Hill

Figure 1.8 Fahrenheit, Celsius, and Kelvin Temperature Scales Compared



Because the freezing point and boiling point of water span 180° on the Fahrenheit scale but only 100° on the Celsius scale, a Fahrenheit degree and a Celsius degree differ in size. The Kelvin scale is divided into kelvins (K), not degrees. Because the freezing point and boiling point of water span 100 kelvins, one kelvin is the same size as one Celsius degree.

100 °C. To convert temperature values from one scale to another, we use two equations, where T_C is the Celsius temperature and T_F is the Fahrenheit temperature.

To convert from Celsius to Fahrenheit:

$$T_F = 1.8(T_C) + 32$$

To convert from Fahrenheit to Celsius:

$$T_C = \frac{T_F - 32}{1.8}$$

STUDY TIP

Whenever a calculation requires a specific equation:

1. Write down the equation.
2. Substitute the known quantities.
3. Solve the equation for the unknown quantity. **If there is more than one unknown, you are using the *wrong* equation.**

The Kelvin scale is divided into **kelvins** (K), not degrees. The only difference between the Kelvin scale and the Celsius scale is the zero point. A temperature of −273 °C corresponds to 0 K. The zero point on the Kelvin scale is called **absolute zero,** the lowest temperature possible. To convert temperature values from Celsius to Kelvin, or vice versa, use two equations.

To convert from Celsius to Kelvin:

$$T_K = T_C + 273$$

To convert from Kelvin to Celsius:

$$T_C = T_K - 273$$

Sample Problem 1.14 Converting Temperature from One Scale to Another

An infant had a temperature of 104 °F. Convert this temperature to both T_C and T_K.

Analysis

First convert the Fahrenheit temperature to degrees Celsius using the equation $T_C = (T_F - 32)/1.8$. Then convert the Celsius temperature to kelvins by adding 273.

Solution

[1] Convert T_F to T_C:

$$T_C = \frac{T_F - 32}{1.8}$$

$$= \frac{104 - 32}{1.8} = 40.\ °C$$

[2] Convert T_C to T_K:

$$T_K = T_C + 273$$

$$= 40. + 273 = 313\ K$$

Practice Problem 1.14

When the human body is exposed to extreme cold, hypothermia can result and the body's temperature can drop to 28.5 °C. Convert this temperature to T_F and T_K.

More Practice: Try Problems 1.53c, 1.54c, 1.55–1.58.

Problem 1.15

Convert each temperature to the requested temperature scale.

a. 20. °C to T_F b. 150. °F to T_C c. 298 K to T_F d. 75 °C to T_K

1.10 Density and Specific Gravity

Two additional quantities used to characterize substances are **density** and **specific gravity.**

1.10A Density

***Density* is a physical property that relates the mass of a substance to its volume.** Density is the ratio of mass (in grams) to volume (in milliliters or cubic centimeters).

$$\text{density} = \frac{\text{mass (g)}}{\text{volume (mL or cc)}}$$

Although a can of a diet soft drink floats in water because it is less dense, a can of a regular soft drink that contains sugar is more dense than water so it sinks.
Jill Braaten/McGraw-Hill

Table 1.6 Representative Densities at 25 °C

Substance	Density [g/(mL or cc)]	Substance	Density [g/(mL or cc)]
Oxygen (0 °C)	0.001 43	Urine	1.003–1.030
Gasoline	0.66	Blood plasma	1.03
Ice (0 °C)	0.92	Table sugar	1.59
Water (4 °C)	1.00	Bone	1.80

The density of a substance depends on temperature. For most substances, the solid state is more dense than the liquid state, and as the temperature increases, the density decreases. This phenomenon occurs because the volume of a sample of a substance generally increases with temperature but the mass is always constant.

Water is an exception to this generalization. Solid water, ice, is *less* dense than liquid water, and from 0 °C to 4 °C, the density of water *increases*. Above 4 °C, water behaves like other liquids and its density decreases. Thus, water's maximum density of 1.00 g/mL occurs at 4 °C. Some representative densities are reported in Table 1.6.

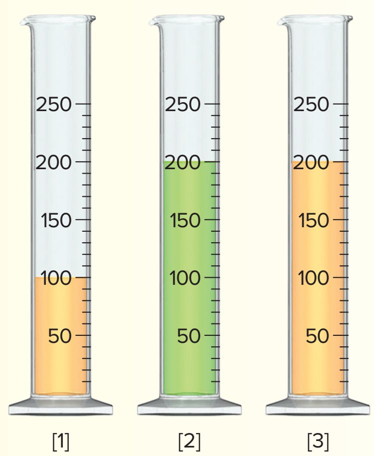
The density (not the mass) of a substance determines whether it floats or sinks in a liquid.

- **A less dense substance floats on a more dense liquid.**

Ice floats on water because it is less dense. When petroleum leaks from an oil tanker or gasoline is spilled when fueling a boat, it floats on water because it is less dense. In contrast, a cannonball or torpedo sinks because it is more dense than water.

Sample Problem 1.15 Using Density to Solve a Problem

The density of liquid **A** is twice the density of liquid **B.** (a) If you have an equal mass of **A** and **B,** which graduated cylinder ([1] or [2]) corresponds to **A** and which corresponds to **B?** (b) How do the masses of the liquids in graduated cylinders [2] and [3] compare?



Analysis

Density is the number of grams per milliliter (g/mL) or grams per cubic centimeter (g/cc) of a substance.

Solution

a. If the density of **A** is twice the density of **B,** you need twice the volume of **B** to have the same mass as a sample of **A.** Thus, graduated cylinder [1] represents **A** (gold liquid) and graduated cylinder [2] represents **B** (green liquid).

b. Because graduated cylinders [2] and [3] have equal volumes of **A** and **B** but **A** is twice as dense as **B,** the mass of [3] (**A**) must be twice the mass of [2] (**B**).