# PRINCIPLES of

## General, Organic, & Biological Chemistry

Third Edition

**Janice Gorzynski Smith**

University of Hawai‘i at Mānoa

Mc Graw Hill

PRINCIPLES OF GENERAL, ORGANIC, & BIOLOGICAL CHEMISTRY, THIRD EDITION

Published by McGraw Hill LLC, 1325 Avenue of the Americas, New York, NY 10019. 

This book is printed on acid-free paper.

1 2 3 4 5 6 7 8 9 LWI 26 25 24 23 22

ISBN 978-1-266-51712-9 (bound edition)
MHID 1-266-51712-X (bound edition)
ISBN 978-1-264-45182-1 (loose-leaf edition)
MHID 1-264-45182-2 (loose-leaf edition)

Associate Portfolio Manager: *Hannah Downing*
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Content Project Managers: *Laura Bies, Rachael Hillebrand*
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Cover Image: *David Merron Photography/Getty Images*
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**Library of Congress Cataloging-in-Publication Data**

Names: Smith, Janice Gorzynski, author.
Title: Principles of general, organic, & biological chemistry / Janice Gorzynski Smith, University of Hawai‘i at Mānoa.
Other titles: Principles of general, organic, and biological chemistry
Description: Third edition. | New York, NY : McGraw Hill Education, [2023] | Includes index.
Identifiers: LCCN 2021030648 (print) | LCCN 2021030649 (ebook) | ISBN 9781266517129 (hardcover) | ISBN 9781264451821 (spiral bound) | ISBN 9781264452385 (pdf)
Subjects: LCSH: Chemistry–Textbooks. | Chemistry, Inorganic–Textbooks. | Biochemistry–Textbooks.
Classification: LCC QD31.3 .S634 2023 (print) | LCC QD31.3 (ebook) | DDC 540–dc23
LC record available at https://lccn.loc.gov/2021030648
LC ebook record available at https://lccn.loc.gov/2021030649



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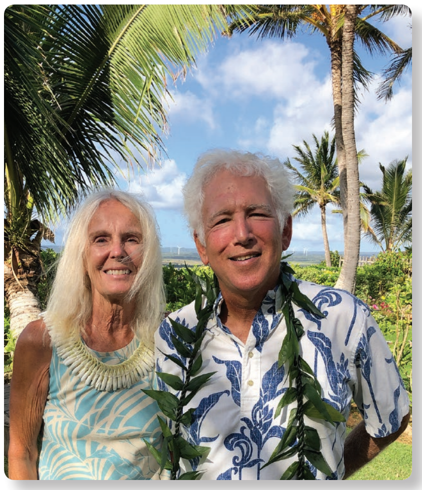
# About the Author



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.

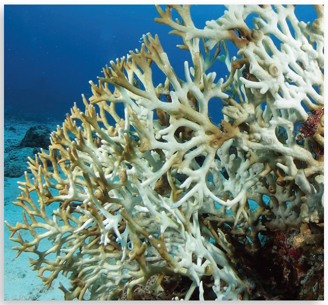
Following her postdoctoral work, Jan joined the faculty of Mount Holyoke College, where she was employed for 21 years. During this time, she was active in teaching organic chemistry lecture and lab courses, conducting a research program in organic synthesis, and serving as department chair. Her organic 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 of Hawai'i in the 1990s, Jan and her family moved there permanently in 2000. She has been a faculty member at the University of Hawai'i at Mānoa, where she has taught a one-semester organic and biological chemistry course for nursing students, as well as the two-semester organic chemistry lecture and lab courses. 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 at an event on Oahu. 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.

*To my family*

# Contents in Brief

# Contents



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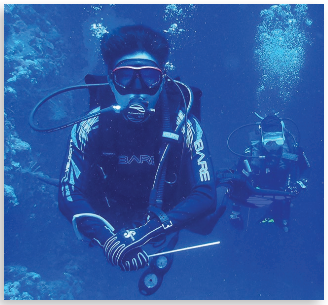

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# Preface

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 one- or two-semester course. This text relates the principal concepts of general, organic, and biological chemistry to the world around us, and in this way illustrates how chemistry explains many aspects of daily life.

The learning style of today's students relies heavily on visual imagery. In this text, new concepts are introduced one at a time, keeping the basic themes in focus, and breaking down complex problems into manageable chunks of information. Relevant applications are provided for all basic chemical concepts, and molecular art illustrates and explains common everyday phenomena. Students learn step-by-step problem solving throughout the chapter within sample problems and *How To* boxes. Students are given enough detail to understand basic concepts so that they may acquire a new appreciation of the human body and the larger world around them.

## New to This Edition

### General

**Problem Solving** Sample Problems are now paired with Practice Problems to allow students to apply what they have just learned. The Practice Problems are followed by More Practice lists that point students to end-of-chapter problems that are similar in concept. Concept Check problems have replaced other in-chapter problems to give students an immediate check on the topic that has just been presented. The answers to all Practice Problems, all Concept Check problems, and odd-numbered end-of-chapter problems are given at the end of each chapter.

**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.

**Self-Test** Each chapter contains a Self-Test, which consists of short-answer questions that test an understanding of definitions, equations, and other material encountered within the chapter. Answers to each question are provided at the end of the chapter.

**Study Tips** Brief Study Tips have been added to the margins in Chapters 1, 3, 6, 10, and 14 to help students develop general methods for solving recurrent types of problems, such as those that require a specific equation.

**Photos** Three-fourths of the chapter-opening photos have been replaced with photos emphasizing relevant material within the chapter. More marginal photos of applications have also been added on topics including non-contact thermometers (Chapter 1), radioactive seeds for cancer treatment (Chapter 9), leghemoglobin in plant-based burgers (Chapter 16), and many others.

**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 8.5B. Tables with art that indicates what buttons should be pressed and what calculator displays will be shown are given.
- Chapter 3 opens by presenting a new current topic, the effect of sunscreens like oxybenzone on the bleaching of coral reefs. Recent research on oxybenzone is also discussed in Section 11.10.
- A new section on determining types of reactions—combination, decomposition, single replacement, and double replacement—has been added to Chapter 5. The chapter has been reorganized to place oxidation and reduction reactions immediately following this section, so that all of these different reaction types are in proximity.
- The discussion of dialysis and osmosis in Section 7.8 has been edited to emphasize the distinction between these related concepts. Three new problems on this subject have been added within the chapter.
- New material on using PET scans to visualize the brain in Alzheimer's patients has been added to Section 9.5.
- Material on methane, a greenhouse gas, has been expanded in Section 10.8.
- New material on the human milk oligosaccharides in breast milk has been added in Section 14.6D.
- Figure 15.1 now presents data on saturated fats, unsaturated oils, and trans fats in bar graph form for easier visualization of lipid content.
- The material on enzymes has been expanded into two sections (Sections 16.9 and 16.10), which include classes of enzymes, naming enzymes, and factors that affect enzyme activity.
- A section on the Human Genome Project (Section 17.10B) has been added.
- Section 17.11 on viruses has been expanded with material on coronaviruses and mRNA vaccines.

## The Construction of a Learning System

Writing a textbook and its supporting learning tools is a multifaceted endeavor. McGraw Hill's development process is an ongoing, market-oriented approach to building accurate and innovative learning systems. It is dedicated to continual large scale and incremental improvement, driven by multiple customer feedback loops and checkpoints. This is initiated during the early planning stages of new products and intensifies during the development and production stages, and then begins again upon publication, in anticipation of the next version of each print and digital product. This process is designed to provide a broad, comprehensive spectrum of feedback for refinement and innovation of learning tools for both student and instructor. The development process includes market research, content reviews, faculty and student focus groups, course- and product-specific symposia, accuracy checks, and art reviews.

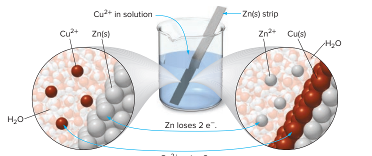
# The Learning System Used in *Principles of General, Organic, & Biological Chemistry,* Third Edition

## Writing Style

A succinct writing style weaves together key points of general, organic, and biological chemistry, along with attention-grabbing applications to consumer, environmental, and health-related fields. Concepts and topics are broken into small chunks of information that are more easily learned.

5.4 Oxidation and Reduction 159

**Figure 5.2** A Redox Reaction—The Transfer of Electrons from Zn to $Cu^{2+}$




A redox reaction occurs when a strip of Zn metal is placed in a solution of $Cu^{2+}$ ions. In this reaction, Zn loses two electrons to form $Zn^{2+}$, which goes into solution. $Cu^{2+}$ gains two electrons to form Cu metal, which precipitates out of solution, forming a coating on the zinc strip.

**CONSUMER NOTE**


Benzoyl peroxide ($C_{14}H_{10}O_4$) is the active ingredient in several acne medications. Benzoyl peroxide kills bacteria by oxidation reactions. *Jill Braaten*

Each of these processes can be written as individual reactions, called **half reactions,** to emphasize which electrons are gained and lost.

Loss of electrons = oxidation

**Oxidation** half reaction: $Zn \longrightarrow Zn^{2+} + \underbrace{2e^-}$

**Reduction** half reaction: $Cu^{2+} + \underbrace{2e^-} \longrightarrow Cu$

Gain of electrons = reduction

- A compound that gains electrons (is reduced) while causing another compound to be oxidized is called an *oxidizing agent.*
- A compound that loses electrons (is oxidized) while causing another compound to be reduced is called a *reducing agent.*

In this example, Zn loses electrons to $Cu^{2+}$. We can think of Zn as a **reducing agent** because it causes $Cu^{2+}$ to gain electrons and become reduced. We can think of $Cu^{2+}$ as an **oxidizing agent** because it causes Zn to lose electrons and become oxidized.

To draw the products of an oxidation–reduction reaction, we must decide which element or ion gains electrons and which element or ion loses electrons. Use the following guidelines.

- When considering neutral atoms, metals *lose* electrons and nonmetals *gain* electrons.
- When considering ions, cations tend to *gain* electrons and anions tend to *lose* electrons.

Thus, the metals sodium (Na) and magnesium (Mg) readily lose electrons to form the cations $Na^+$ and $Mg^{2+}$, respectively; that is, they are oxidized. The nonmetals $O_2$ and $Cl_2$ readily gain electrons to form 2 $O^{2-}$ and 2 $Cl^-$, respectively; that is, they are reduced. A positively charged ion like $Cu^{2+}$ is reduced to Cu by gaining two electrons, whereas two negatively charged $Cl^-$ anions are oxidized to $Cl_2$ by losing two electrons. These reactions and additional examples are shown in Figure 5.3.

## Chapter Goals, Tied to End-of-Chapter Review

Chapter Goals at the beginning of each chapter identify what students will learn, and are tied to the end-of-chapter Key Concepts and Key Skills, which serve as bulleted summaries of the most important concepts for study.

**CHAPTER OUTLINE**

**9.1** Isotopes and Radioactivity
**9.2** Nuclear Reactions
**9.3** Half-Life
**9.4** Detecting and Measuring Radioactivity
**9.5** FOCUS ON HEALTH & MEDICINE: Medical Uses of Radioisotopes
**9.6** Nuclear Fission and Nuclear Fusion
**9.7** FOCUS ON HEALTH & MEDICINE: Medical Imaging Without Radioactivity

**CHAPTER GOALS**

*In this chapter you will learn how to:*

1. Describe the different types of radiation emitted by a radioactive nucleus
2. Write equations for nuclear reactions
3. Define half-life
4. Recognize the units used for measuring radioactivity
5. Give examples of common radioisotopes used in medical diagnosis and treatment
6. Describe the general features of nuclear fission and nuclear fusion
7. Describe the features of medical imaging techniques that do not use radioactivity

**KEY CONCEPTS**

**1 Types of nuclear radiation (9.1)**

| 1 Alpha particle | 2 Beta particle | 3 Positron | 4 Gamma ray |
|---|---|---|---|
| $\alpha$ or ${}^{4}_{2}He$ | $\beta$ or ${}^{0}_{-1}e$ | $\beta^+$ or ${}^{0}_{+1}e$ | $\gamma$ |
| • A high-energy nucleus that contains two protons and two neutrons | • A high-energy electron that has a −1 charge and negligible mass | • An antiparticle of a β particle that has a +1 charge and negligible mass | • High-energy radiation with no mass or charge |

**2 Nuclear fission and nuclear fusion (9.6)**

| 1 Nuclear fission | 2 Nuclear fusion |
|---|---|
| ${}^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{91}_{36}Kr + {}^{142}_{56}Ba + 3\,{}^{1}_{0}n$ | ${}^{2}_{1}H + {}^{3}_{1}H \longrightarrow {}^{4}_{2}He + {}^{1}_{0}n$ |
| • Nuclear fission is the splitting apart of a heavy nucleus into lighter nuclei and neutrons.<br>• Nuclear fission releases a great deal of energy.<br>• Fission is used in nuclear power plants to generate electricity. | • Nuclear fusion is the joining together of two light nuclei to form a larger nucleus.<br>• Nuclear fusion releases a great deal of energy.<br>• Nuclear fusion occurs in stars. |



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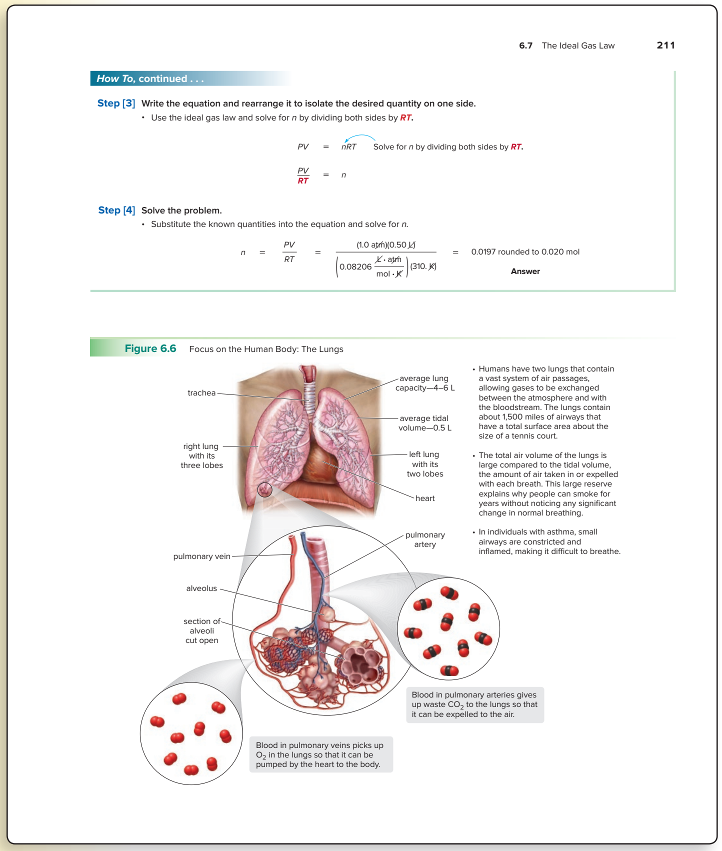
## Macro-to-Micro Illustrations

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 visualize and understand the chemistry behind ordinary things. Many illustrations of the human body include magnifications for specific anatomic regions, as well as representations at the microscopic level, for today's visual learners.



6.7 The Ideal Gas Law 211

*How To*, continued . . .

**Step [3]** **Write the equation and rearrange it to isolate the desired quantity on one side.**

- Use the ideal gas law and solve for *n* by dividing both sides by ***RT***.

$$PV = nRT \quad \text{Solve for } n \text{ by dividing both sides by } \mathbf{RT}.$$

$$\frac{PV}{RT} = n$$

**Step [4]** **Solve the problem.**

- Substitute the known quantities into the equation and solve for *n*.

$$n = \frac{PV}{RT} = \frac{(1.0\ \cancel{\text{atm}})(0.50\ \cancel{\text{L}})}{\left(0.08206\ \frac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\text{mol}\cdot\cancel{\text{K}}}\right)(310.\ \cancel{\text{K}})} = 0.0197 \text{ rounded to } 0.020 \text{ mol}$$

**Answer**

Figure 6.6 Focus on the Human Body: The Lungs



- Humans have two lungs that contain a vast system of air passages, allowing gases to be exchanged between the atmosphere and with the bloodstream. The lungs contain about 1,500 miles of airways that have a total surface area about the size of a tennis court.
- The total air volume of the lungs is large compared to the tidal volume, the amount of air taken in or expelled with each breath. This large reserve explains why people can smoke for years without noticing any significant change in normal breathing.
- In individuals with asthma, small airways are constricted and inflamed, making it difficult to breathe.

## Applications

Relevant, interesting applications of chemistry to everyday life are included for all basic chemical concepts. These are interspersed in margin-placed Health Notes, Consumer Notes, and Environmental Notes, as well as sections entitled "Focus on Health & Medicine," "Focus on the Environment," and "Focus on the Human Body."

8.9 Focus on the Human Body: Buffers in the Blood 289

HEALTH NOTE



Individuals with cystic fibrosis, the most common genetic disease in Caucasians, produce thick mucus in the lungs, resulting in a higher-than-normal level of $CO_2$ and respiratory acidosis.

HEALTH NOTE



During strenuous exercise, the lungs expel more $CO_2$ than usual and the pH of the blood increases.
*CS Productions/Jupiterimages/Brand X Pictures/Alamy Stock Photo*

### 8.9 FOCUS ON THE HUMAN BODY Buffers in the Blood

The normal blood pH of a healthy individual is in the range of 7.35 to 7.45. A pH above or below this range is generally indicative of an imbalance in respiratory or metabolic processes. The body is able to maintain a very stable pH because the blood and other tissues are buffered. The principal buffer in the blood is carbonic acid/bicarbonate ($H_2CO_3/HCO_3^-$).

In examining the carbonic acid/bicarbonate buffer system in the blood, two reactions are important. First of all, carbonic acid ($H_2CO_3$) is formed from $CO_2$ dissolved in the bloodstream (Section 8.6). Second, because carbonic acid is a weak acid, it is also dissociated in water to form its conjugate base, bicarbonate ($HCO_3^-$). Bicarbonate is also generated in the kidneys.

$$CO_2(g) + H_2O(l) \rightleftharpoons \underset{\text{carbonic acid}}{H_2CO_3(aq)} \overset{H_2O}{\rightleftharpoons} H_3O^+(aq) + \underset{\text{bicarbonate}}{HCO_3^-(aq)}$$

principal buffer in the blood

$CO_2$ is constantly produced by metabolic processes in the body and then transported to the lungs to be eliminated. Le Châtelier's principle explains the effect of increasing or decreasing the level of dissolved $CO_2$ on the pH of the blood. A higher-than-normal $CO_2$ concentration shifts the equilibrium to the right, increasing the $H_3O^+$ concentration and lowering the pH. **Respiratory acidosis** results when the body fails to eliminate adequate amounts of $CO_2$ through the lungs. This may occur in patients with advanced lung disease or respiratory failure.

A lower respiratory rate **increases [$CO_2$]**.

$$CO_2(g) + 2\,H_2O(g) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

This drives the reaction to the right, **increasing [$H_3O^+$]**.
Blood has a higher [$H_3O^+$] $\longrightarrow$ **lower pH**

A lower-than-normal $CO_2$ concentration shifts the equilibrium to the left, decreasing the $H_3O^+$ concentration and raising the pH. **Respiratory alkalosis** is caused by hyperventilation, very rapid breathing that occurs when an individual experiences excitement or panic.

$$CO_2(g) + 2\,H_2O(g) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

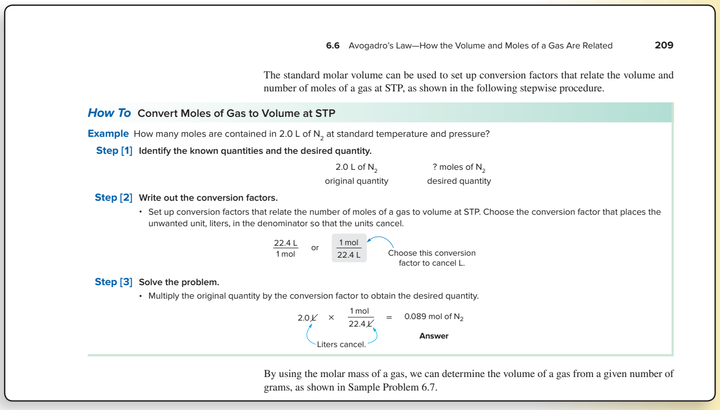
A faster respiratory rate **decreases [$CO_2$]**.

This drives the reaction to the left, **decreasing [$H_3O^+$]**.
Blood has a lower [$H_3O^+$] $\longrightarrow$ **higher pH**

The pH of the blood may also be altered when the metabolic processes of the body are not in balance. **Metabolic acidosis** results when excessive amounts of acid are produced and the blood pH falls. This may be observed in patients with severe infections (sepsis). It may also occur in poorly controlled diabetes. **Metabolic alkalosis** may occur when recurrent vomiting decreases the amount of acid in the stomach, thus causing a rise in pH.

## *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.



6.6 Avogadro's Law—How the Volume and Moles of a Gas Are Related 209

The standard molar volume can be used to set up conversion factors that relate the volume and number of moles of a gas at STP, as shown in the following stepwise procedure.

***How To*** **Convert Moles of Gas to Volume at STP**

**Example** How many moles are contained in 2.0 L of $N_2$ at standard temperature and pressure?

**Step [1] Identify the known quantities and the desired quantity.**

2.0 L of $N_2$ (original quantity) ? moles of $N_2$ (desired quantity)

**Step [2] Write out the conversion factors.**

- Set up conversion factors that relate the number of moles of a gas to volume at STP. Choose the conversion factor that places the unwanted unit, liters, in the denominator so that the units cancel.

$$\frac{22.4\ \text{L}}{1\ \text{mol}} \quad \text{or} \quad \boxed{\frac{1\ \text{mol}}{22.4\ \text{L}}}$$

Choose this conversion factor to cancel L.

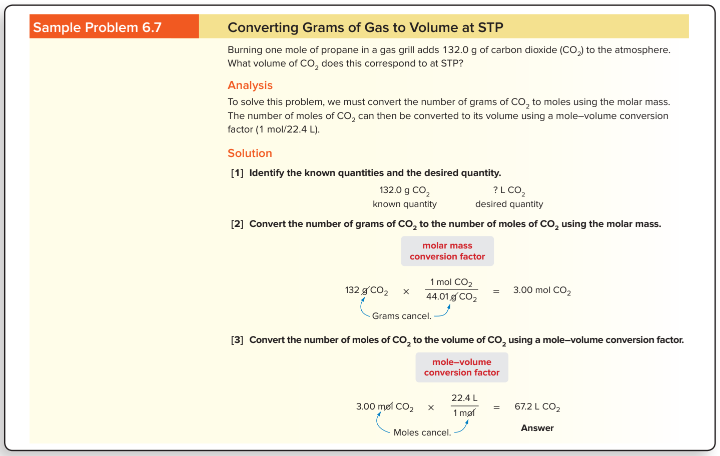
**Step [3] Solve the problem.**

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

$$2.0\ \cancel{\text{L}} \times \frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}} = 0.089\ \text{mol of } N_2$$

Liters cancel. **Answer**

By using the molar mass of a gas, we can determine the volume of a gas from a given number of grams, as shown in Sample Problem 6.7.



**Sample Problem 6.7 Converting Grams of Gas to Volume at STP**

Burning one mole of propane in a gas grill adds 132.0 g of carbon dioxide ($CO_2$) to the atmosphere. What volume of $CO_2$ does this correspond to at STP?

**Analysis**

To solve this problem, we must convert the number of grams of $CO_2$ to moles using the molar mass. The number of moles of $CO_2$ can then be converted to its volume using a mole–volume conversion factor (1 mol/22.4 L).

**Solution**

**[1] Identify the known quantities and the desired quantity.**

132.0 g $CO_2$ (known quantity) ? L $CO_2$ (desired quantity)

**[2] Convert the number of grams of $CO_2$ to the number of moles of $CO_2$ using the molar mass.**

molar mass conversion factor

$$132\ \cancel{\text{g}}\ CO_2 \times \frac{1\ \text{mol } CO_2}{44.01\ \cancel{\text{g}}\ CO_2} = 3.00\ \text{mol } CO_2$$

Grams cancel.

**[3] Convert the number of moles of $CO_2$ to the volume of $CO_2$ using a mole–volume conversion factor.**

mole–volume conversion factor

$$3.00\ \cancel{\text{mol}}\ CO_2 \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 67.2\ \text{L } CO_2$$

Moles cancel. **Answer**

## Problem Solving

Stepwise sample problems lead students through the thought process tied to successful problem solving by employing *Analysis* and *Solution* steps. Sample Problems are categorized sequentially by topic to match chapter organization, and are paired with practice problems to allow students to apply what they have just learned. Students can immediately verify their answers to the follow-up problems in the answers at the end of each chapter.

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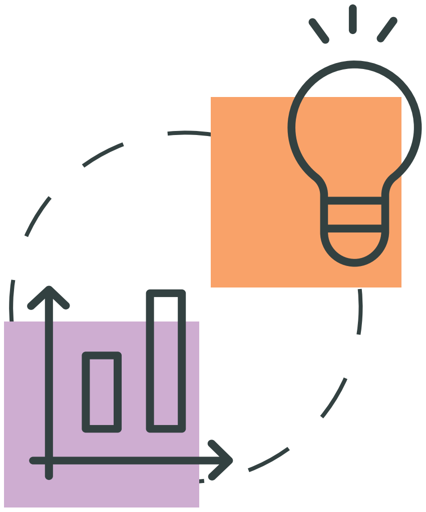

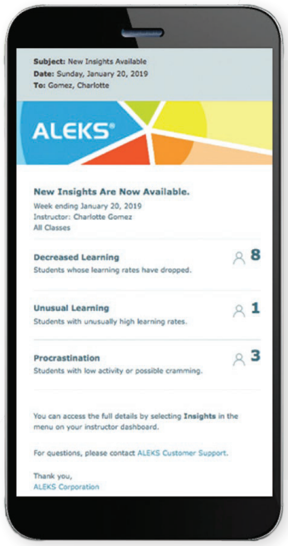


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- **Presentation Tools Table:** Instructors have access to fully editable **accessible PowerPoint lecture outlines,** which appear as ready-made presentations that combine art and lecture notes for each chapter of the text. For instructors who prefer to create their lectures from scratch, all illustrations, photos, and tables are pre-inserted by chapter into blank **PowerPoint slides** and are also available as **downloadable jpeg files.**
- **Instructor's Solutions Manual:** This supplement contains complete, worked out solutions for all the end-of-chapter problems in the text.
- **Computerized Test Bank:** Over 1,800 test questions that accompany *Principles of General, Organic, & Biological Chemistry* are available for creating exams or quizzes.
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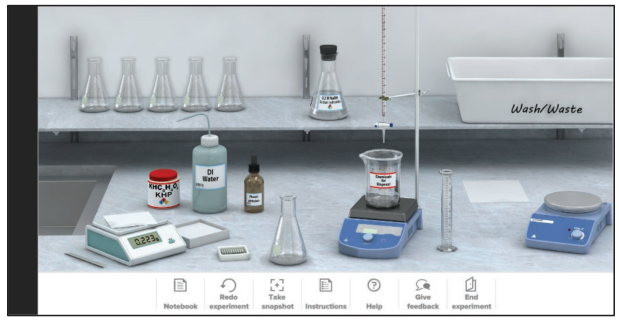


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## Student Solutions Manual

The *Student Solutions Manual* contains the solutions to all in-chapter problems as well as the solutions to all odd-numbered end-of-chapter problems.

# Acknowledgments

When I first began textbook writing 20 years ago, I had no idea how many people I would have to rely upon to see a project from manuscript preparation to published text. Special thanks for this edition go to Senior Product Developer Mary Hurley, Senior Core Content Project Manager Laura Bies, and freelance Developmental Editor John Murdzek, who handled all the day-to-day steps needed to publish an accurate and engaging text. Thanks are also due to Executive Portfolio Manager Michelle Hentz and Associate Portfolio Manager Hannah Downing for spearheading the revision of *Principles,* and the art, production, marketing, and sales teams for their support and contributions.

I would also like to acknowledge the following individuals for their masterful authoring of the ancillaries to accompany the third edition: Lauren McMills of Ohio University for the Solutions Manuals; Andrea Leonard of the University of Louisiana, Lafayette for the Accessible PowerPoint Lecture Outlines; and Cari Gigliotti of Sinclair Community College for the Test Bank.

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, took several photos that appear in the text, and served as a consultant for many medical applications.

# 1 Matter and Measurement


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, 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, showered 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 or 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 a Natural Material into a Useful Synthetic Product

a.

b.

(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 to 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 to 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 DVDs, 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.

## 1.2 States of Matter

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

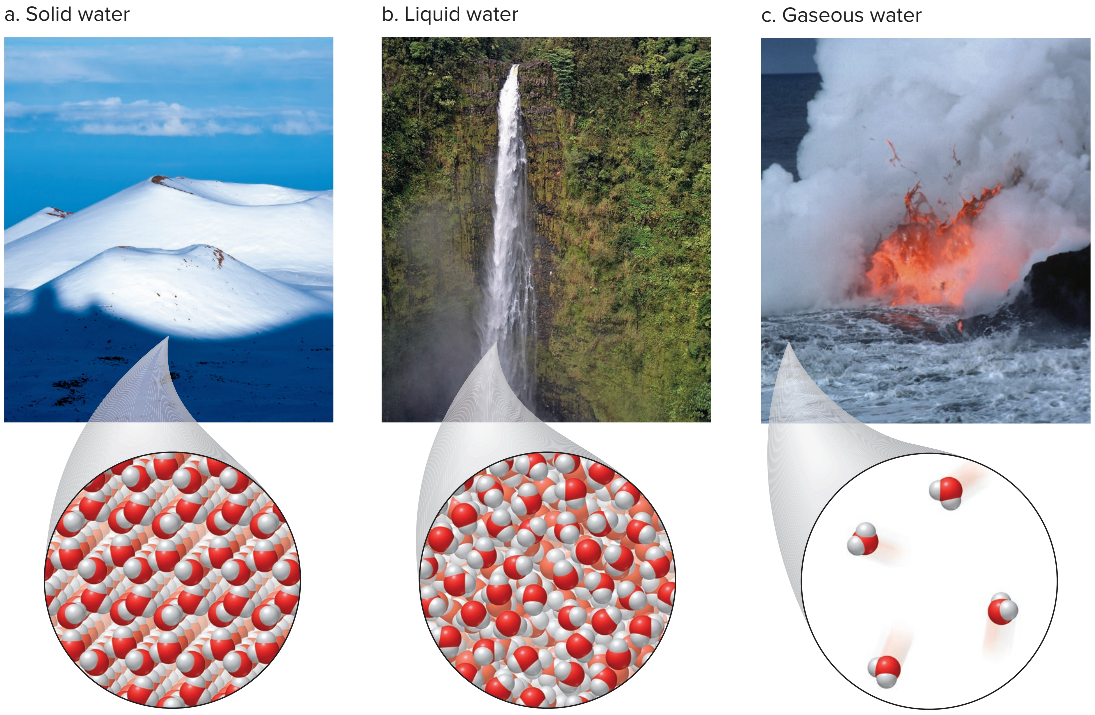
- A *solid* has a definite volume, and maintains its shape regardless of the container in which it is placed. The particles of a solid lie close together, and are arranged in a regular three-dimensional array.
- A *liquid* has a definite volume, but takes on the shape of the container it occupies. The 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.

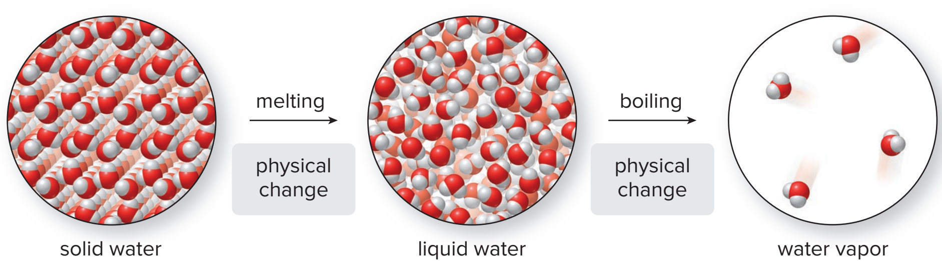
**Figure 1.3** The Three States of Water—Solid, Liquid, and Gas




- 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)

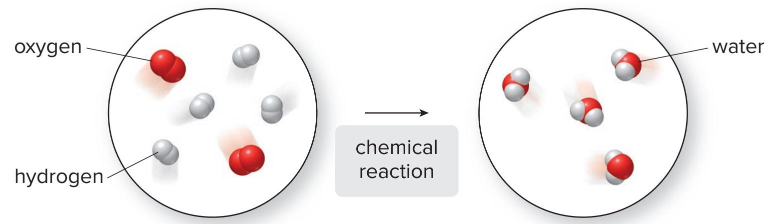
Each red sphere joined to two gray spheres represents a single water particle. In proceeding from left to right, from solid to liquid to gas, the molecular art shows that the level of organization of the water particles decreases. Color-coding and the identity of the spheres within the particles will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

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 4.




- ***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 to another.** The conversion of hydrogen and oxygen to 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 Chapter 5.

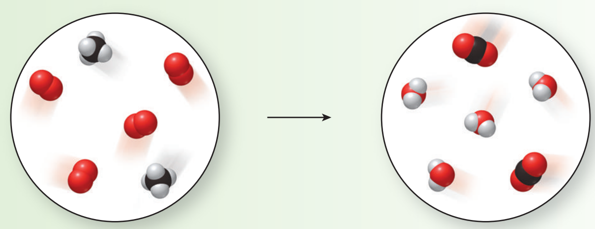



**Concept Check 1.1**

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.

**Concept Check 1.2**

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



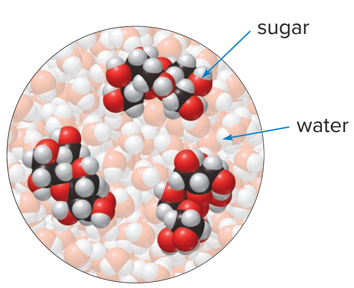
## 1.3 Classification of Matter

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.**

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.**



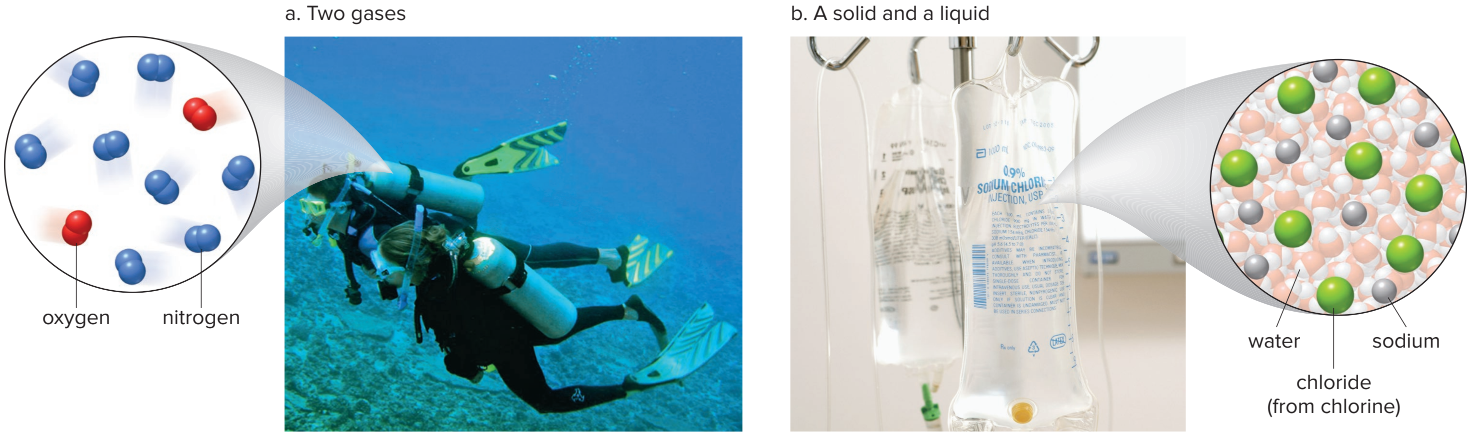
sugar dissolved in water

mixture

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.

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 water.

**Figure 1.4** Two Examples of Mixtures




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

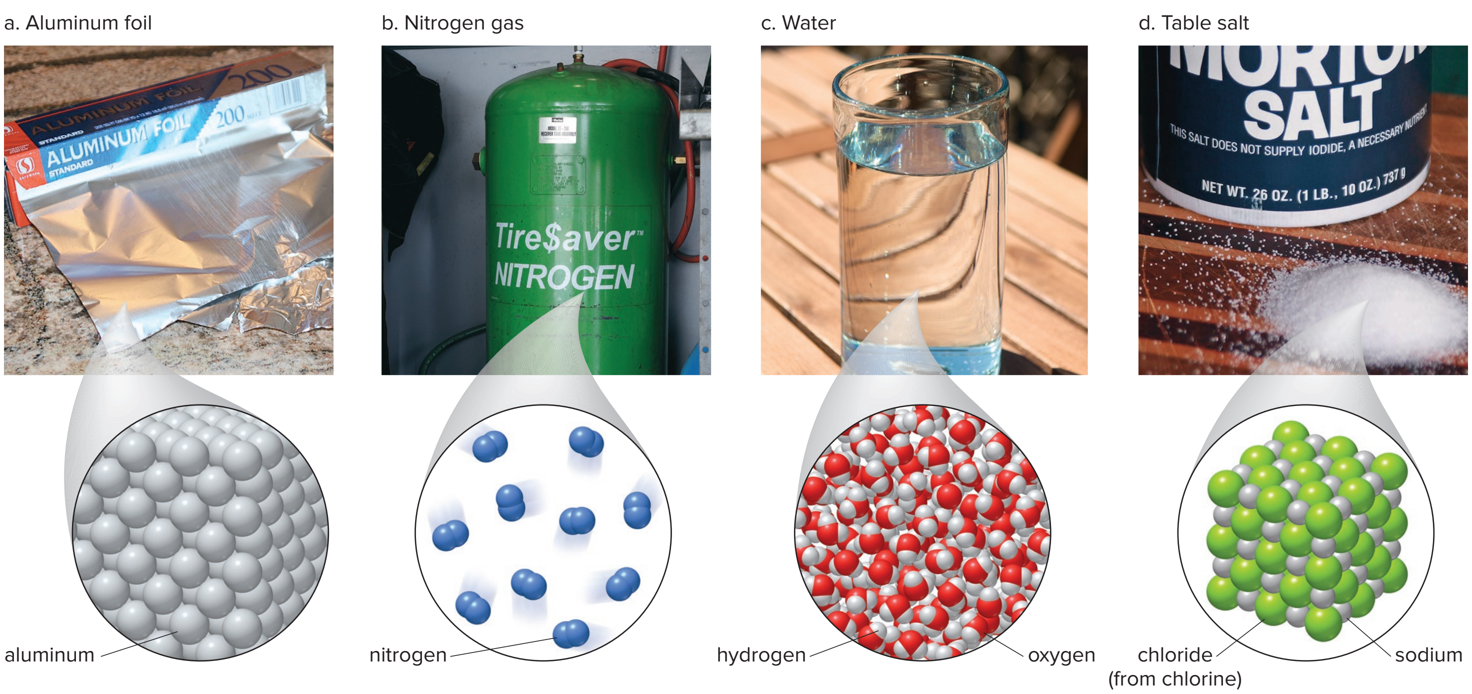
A pure substance is classified as either an **element** or a **compound.**

- 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 Chapter 2.

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

**Figure 1.5** Elements and Compounds

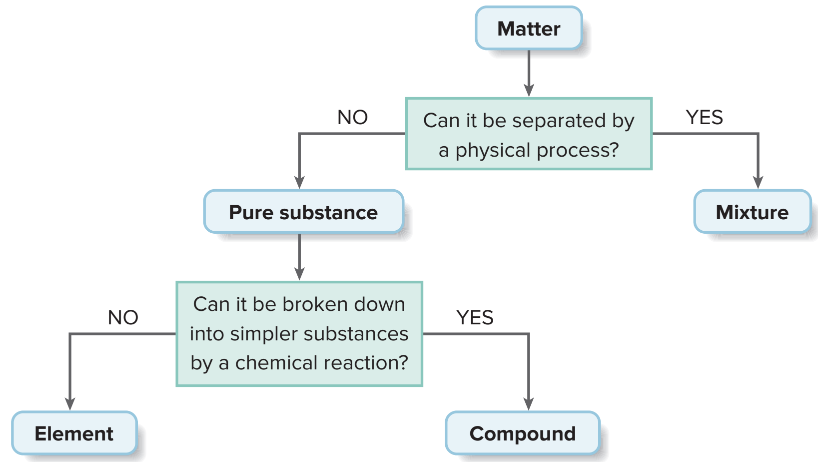



- Aluminum foil and nitrogen gas are elements. **The molecular art used for an element shows spheres of *one* color only.** Thus, aluminum is a solid shown with gray spheres, whereas nitrogen is a gas shown with blue spheres. Water and table salt are compounds. Color-coding of the spheres used in the 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




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 Chapters 2 and 3.

**Concept Check 1.3**

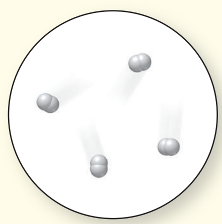
Use the list of elements in Appendix A to classify each item as an element or a compound: (a) the gas inside a helium balloon; (b) table sugar; (c) the rust on an iron nail; (d) aspirin.

Figure 1.6 summarizes the categories into which matter is classified.

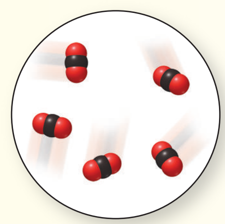
## Sample Problem 1.1 Using Molecular Art for an Element and a Compound

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.

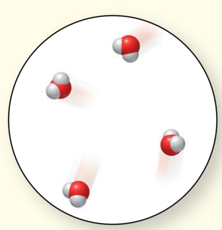
### Solution

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.

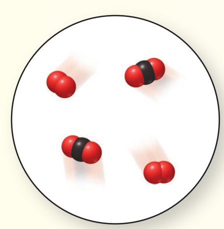
**Practice Problem 1.1**

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

a.



b.



**Practice Problem 1.2**

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.

More Practice: Try Problems 1.1, 1.2, 1.15, 1.16.

## 1.4 Measurement

Any time 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.

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.*

### 1.4A 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.

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.**

**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*

**Table 1.1 Metric Units**

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

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** meters or 1 mm = 0.001 m
1 **milli**gram = **0.001** grams or 1 mg = 0.001 g
1 **milli**liter = **0.001** liters or 1 mL = 0.001 L

**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 lower case Greek letter mu. The prefix *micro-* is sometimes abbreviated as **mc.**

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

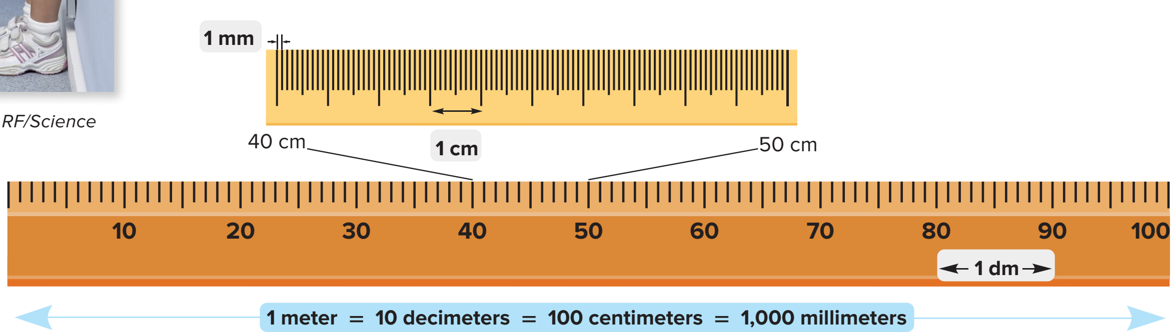
*Adam Gault/Science Photo Library RF/Science Source*

**Concept Check 1.4**

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?

## 1.4B Measuring Length

**The base unit of length in the metric system is the *meter* (m).** A meter, 39.37 inches in the English system, is slightly longer than a yard (36 inches). Common units derived from a meter 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.

*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 = 453.6 g). Two common units derived from a gram are the kilogram (kg) and milligram (mg).

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

**Concept Check 1.5**

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

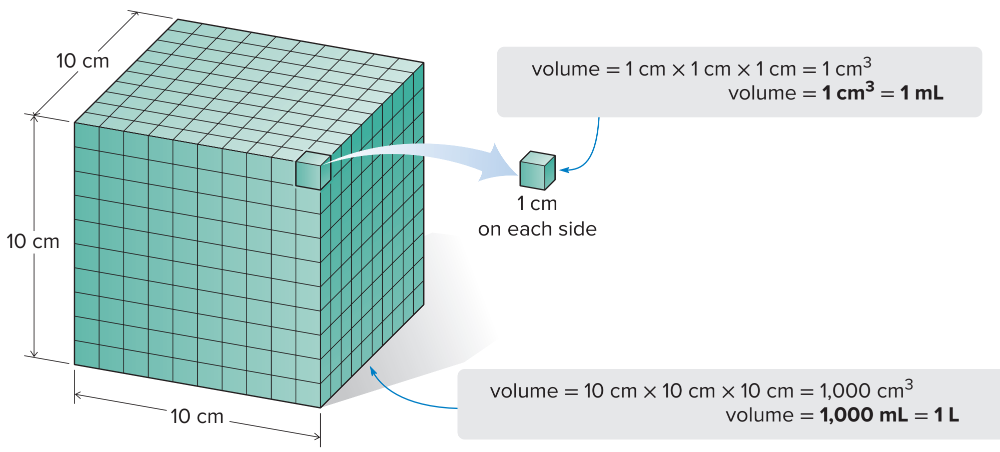
Note the difference between the units **cm** and **cm³.** The centimeter (cm) is a unit of length. A cubic centimeter ($cm^3$ or cc) is a unit of volume.

*Rawpixel/123RF*

$1 \text{ mL} = 1 \text{ cm}^3 = 1 \text{ cc}$

## 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 = 1.057 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.**

$$1 \text{ L} = 10 \text{ dL}$$
$$1 \text{ L} = 1{,}000 \text{ mL}$$
$$1 \text{ L} = 1{,}000{,}000 \text{ μL}$$

**Concept Check 1.6**

If an IV bag contains 1,000 mL of solution, how many liters does it contain?

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.

**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 = 39.37 in. |
| | 1 mi = 5,280. ft | 1 km = 0.6214 mi |
| Mass | 1 lb = 16.00 oz | 1 kg = 2.205 lb |
| | 1 ton = 2,000 lb | 453.6 g = 1 lb |
| | | 28.35 g = 1 oz |
| Volume | 1 qt = 4 cups | 946.4 mL = 1 qt |
| | 1 qt = 2 pt | 1 L = 1.057 qt |
| | 1 qt = 32 fl oz | 29.57 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).

**Concept Check 1.7**

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

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

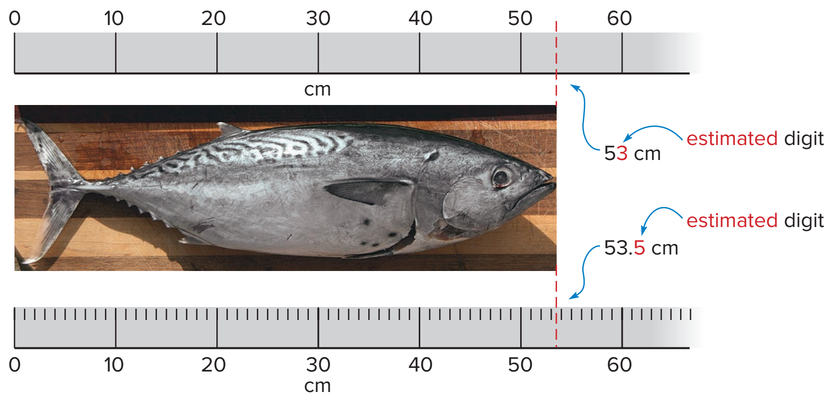
Numbers used in chemistry are either **exact** or **inexact.**

- **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 1.2 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 only if 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 b. 230 c. 0.202 d. 0.003 60 e. 10,040 f. 1,004.00

**Practice Problem 1.4**

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

More Practice: Try Problems 1.5b, 1.6b, 1.23, 1.24.

## 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.**8**54 545

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 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.

Table 1.5 gives other examples of rounding off numbers.

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 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** |

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 8.5B).
*S.Narongrit/Shutterstock*

## Sample Problem 1.3 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 nonzero 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 1.5**

Round off each number in Sample Problem 1.3 to two significant figures.

**More Practice:** Try Problems 1.25, 1.26.

## Sample Problem 1.4 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 ÷ 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 fewest number of significant figures.

**Solution**

a. 3.81 × 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 1.6**

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 1.27a, c; 1.28a, c.

## 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 1.5 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 1.7

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 1.27b, d; 1.28b, d.

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

## 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.**

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.

**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. Because 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. Because 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 = $3.6 \times 10^{-2}$**

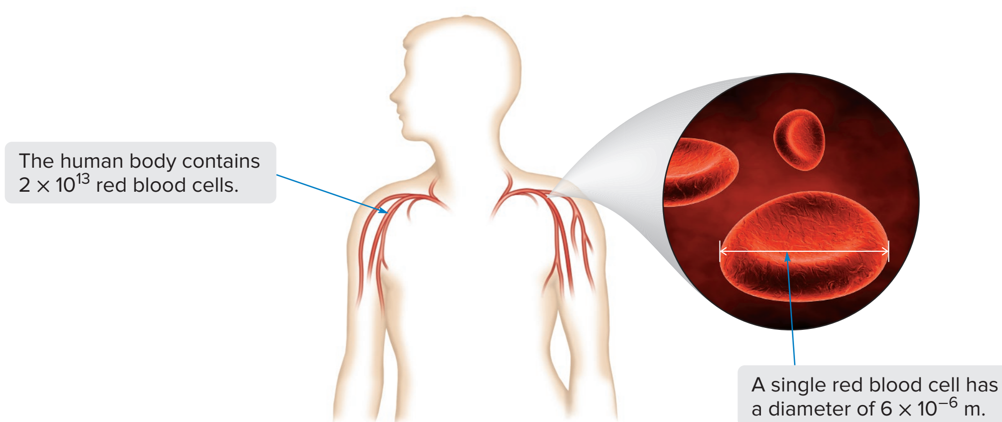
Notice that the number of significant figures in the coefficient in scientific notation must equal the number of significant figures in the original number. Thus, the coefficients for both 2,500 and 0.036 need two significant figures and no more. Figure 1.7 shows two more examples of numbers written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

$\underbrace{2{,}5}_{\text{two significant figures}}00 = 2.5 \times 10^3$ ***not*** $2.50 \times 10^3$ (three significant figures)

***not*** $2.500 \times 10^3$ (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 | $2 \times 10^{13}$ |
| Diameter of a red blood cell | 0.000 006 m | $6 \times 10^{-6}$ m |

## Sample Problem 1.6 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 × $10^3$ — 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 × $10^{-6}$ — 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 1.8

Lab results for a routine check-up showed an individual's iron level in the blood to be 0.000 098 g per deciliter, placing it in the normal range. Convert this number to scientific notation.

## Practice Problem 1.9

Write each number in scientific notation.

a. 93,200 b. 0.000 725 c. 6,780,000 d. 0.000 030

More Practice: Try Problems 1.29, 1.30, 1.35, 1.36.

To convert a number in scientific notation to a standard number, reverse the procedure, as shown in Sample Problem 1.7. 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 1.7 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 1.10**

Convert each number to its standard form.

a. $6.5 \times 10^{3}$ b. $3.26 \times 10^{-5}$ c. $3.780 \times 10^{-2}$ d. $1.04 \times 10^{8}$

More Practice: Try Problems 1.31, 1.32.

### 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 | $1.2^{03}$ | $1.2 \times 10^{3}$ |
| A **decimal** < 1: | 0.052 | 2nd SCI | $5.2^{-02}$ | $5.2 \times 10^{-2}$ |

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 $1.5 \times 10^{8}$: | 1.5 | EE 8 | $1.5^{08}$ | 150,000,000 |
| A number with a **negative** exponent, as in $3.5 \times 10^{-4}$: | 3.5 | EE 4 Change sign (+ to –) | $3.5^{-04}$ | 0.000 35 |

**Concept Check 1.8**

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.

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.

A conversion factor is formed by taking an equality, such as 2.205 lb = 1 kg, and writing it as a fraction. We can always write a conversion factor in two different ways.

$$\frac{2.205\ \text{lb}}{1\ \text{kg}} \quad \text{or} \quad \frac{1\ \text{kg}}{2.205\ \text{lb}}$$

numerator

denominator

**conversion factors** for pounds and kilograms

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.

**Concept Check 1.9**

Write two conversion factors for each pair of units: (a) miles and kilometers; (b) meters and millimeters.

## 1.7B Solving a Problem Using One Conversion Factor

When using conversion factors to solve a problem, a unit that appears in the numerator in one term and the denominator in another term will *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.

| 130 lb | × | conversion factor | = | ? kg |
|---|---|---|---|---|
| original quantity | | | | desired quantity |

Two possible conversion factors: $\frac{2.205\ \text{lb}}{1\ \text{kg}}$ or $\frac{1\ \text{kg}}{2.205\ \text{lb}}$


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

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

- **The conversion factor must relate 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

$$130\ \cancel{\text{lb}} \times \frac{1\ \text{kg}}{2.205\ \cancel{\text{lb}}} = 59\ \text{kg}$$

**answer in kilograms**

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

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 2.205 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. Because 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). Because 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.

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

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 1.8 Solving a Problem Using Conversion Factors

**ENVIRONMENTAL NOTE**

As we will learn in Chapter 11, PET, the plastic used in water bottles, can be recycled into bags, containers, fleece jackets, and carpeting. *Angela Hampton Picture Library/Alamy Stock Photo*

If $1.65 \times 10^9$ 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?

### Analysis and Solution

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

| $1.65 \times 10^9$ 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.

$\dfrac{2.205\text{ lb}}{1\text{ kg}}$ or $\boxed{\dfrac{1\text{ kg}}{2.205\text{ lb}}}$ 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.

$$1.65 \times 10^9\ \cancel{\text{lb}} \times \underbrace{\frac{1\ \text{kg}}{2.205\ \cancel{\text{lb}}}}_{\text{conversion factor}} = 7.48 \times 10^8\ \text{kg of PET}$$

The number of lb (unwanted unit) cancels.

**Practice Problem 1.11**

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

**Practice Problem 1.12**

If the mass of the baby in the chapter-opening photograph is 6 lb 12 oz, how many kilograms does the baby weigh?

**Practice Problem 1.13**

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

More Practice: Try Problems 1.10; 1.37a, c; 1.38b, c; 1.39; 1.40; 1.41a; 1.42a, b; 1.44; 1.57.

## 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 1.9 illustrates how to solve a problem with two conversion factors.

### Sample Problem 1.9 Solving a Problem Using More Than One Conversion Factor

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

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

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

quart–liter conversion: $\frac{1.057\ \text{qt}}{1\ \text{L}}$ or $\boxed{\frac{1\ \text{L}}{1.057\ \text{qt}}}$

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 of 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.

$$1.0\ \cancel{\text{pt}} \times \frac{1\ \cancel{\text{qt}}}{2\ \cancel{\text{pt}}} \times \frac{1\ \text{L}}{1.057\ \cancel{\text{qt}}} = 0.47\ \text{L}$$

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 1.14

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 1.9; 1.14; 1.37b, d; 1.38a, d; 1.41b; 1.43.

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

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 1.10 and 1.11 illustrate how these conversion factors are used in determining drug dosages.

**HEALTH NOTE**


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

### Sample Problem 1.10 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. Because 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 1.15**

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 Problems 1.13, 1.59.

## Sample Problem 1.11 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.


*Mark Dierker/McGraw Hill*

**[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.

**Practice Problem 1.16**

(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?

**Practice Problem 1.17**

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?

More Practice: Try Problem 1.60.

**HEALTH NOTE**

To control the spread of contagious diseases, the body temperature of hospital visitors, workplace employees, and airline travelers is now often measured using a non-contact forehead thermometer.
*Batechenkofff/Shutterstock*

## 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 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.

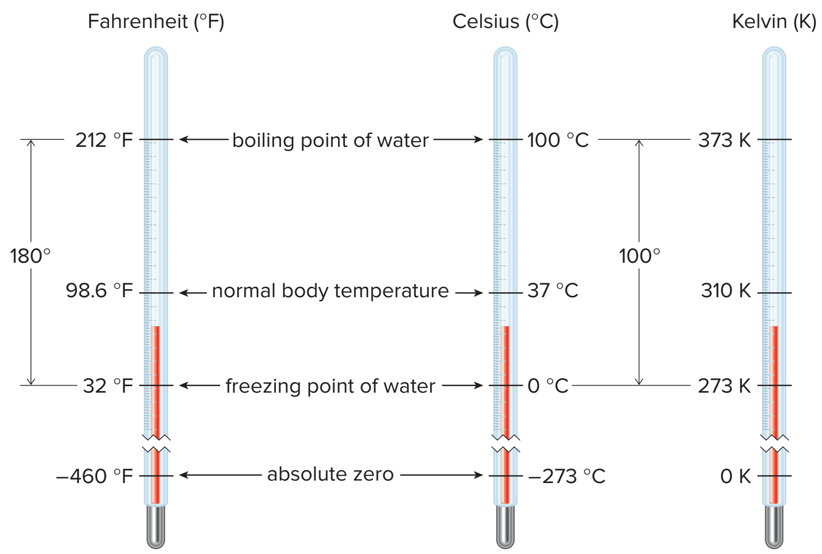
To convert from Celsius to Fahrenheit:

$$T_F = \frac{9}{5}(T_C) + 32$$

To convert from Fahrenheit to Celsius:

$$T_C = \frac{5}{9}(T_F - 32)$$

**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.

**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:

$$T_K = T_C + 273.15$$

To convert from Kelvin to Celsius:

$$T_C = T_K - 273.15$$

## Sample Problem 1.12 Converting Temperature from One Scale to Another

An infant had a temperature of 104 °F. Convert this temperature to both °C and K.

### Analysis

First convert the Fahrenheit temperature to degrees Celsius using the equation $T_C = 5/9(T_F - 32)$. Then convert the Celsius temperature to kelvins by adding 273.15.

### Solution

**[1] Convert $T_F$ to $T_C$:**

$$T_C = \frac{5}{9}(T_F - 32)$$

$$= \frac{5}{9}(104 - 32) = 40.\ °C$$

**[2] Convert $T_C$ to $T_K$:**

$$T_K = T_C + 273.15$$

$$= 40. + 273.15 = 313\ K$$

**Practice Problem 1.18**

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 $T_K$.

**Practice Problem 1.19**

Convert each temperature to the requested temperature scale.

a. 20. °C to $T_F$  b. 150 °F to $T_C$  c. 75 °C to $T_K$

More Practice: Try Problems 1.5c, 1.41c, 1.42c, 1.45.

## 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 reported in grams per milliliter (g/mL) or grams per cubic centimeter (g/cc).

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

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.

**Table 1.6 Representative Densities at 25 °C**

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

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*

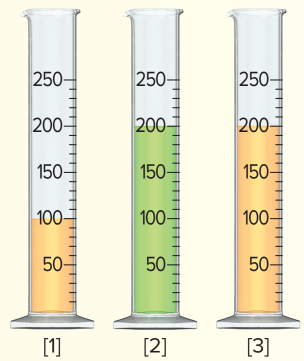
The density (not the mass) of a substance determines whether it floats or sinks in a liquid.

- **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 1.13 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.**

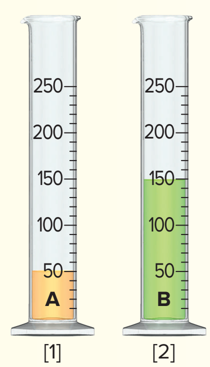
### Solution

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**).

## Practice Problem 1.20

How does the mass of liquid **A** in cylinder [1] compare with the mass of liquid **B** in cylinder [2] in each case (greater than, less than, or equal to)? (a) The densities of **A** and **B** are the same. (b) The density of **A** is twice the density of **B.** (c) The density of **B** is twice the density of **A.**




**More Practice:** Try Problem 1.50.

Knowing the density of a liquid allows us to convert the volume of a substance to its mass, or the mass of a substance to its volume.

**To convert volume (mL) to mass (g):**

$$\mathrm{\cancel{mL}} \times \underbrace{\frac{\mathrm{g}}{\mathrm{\cancel{mL}}}}_{\text{density}} = \mathrm{g}$$

Milliliters cancel.

**To convert mass (g) to volume (mL):**

$$\mathrm{\cancel{g}} \times \underbrace{\frac{\mathrm{mL}}{\mathrm{\cancel{g}}}}_{\text{inverse of the density}} = \mathrm{mL}$$

Grams cancel.

For example, one laboratory synthesis of aspirin uses the liquid acetic acid, which has a density of 1.05 g/mL. If we need 5.0 g for a synthesis, we could use density to convert this mass to a volume that could then be easily measured out using a syringe or pipette.

$$5.0\ \cancel{\mathrm{g}}\ \text{acetic acid} \times \frac{1\ \mathrm{mL}}{1.05\ \cancel{\mathrm{g}}} = 4.8\ \text{mL of acetic acid}$$

Grams cancel.

## Sample Problem 1.14 Using Density and Volume to Determine Mass

Calculate the mass in grams of 15.0 mL of a saline solution that has a density of 1.05 g/mL.

### Analysis

Use density (g/mL) to interconvert the mass and volume of a liquid.

### Solution

$$15.0\ \cancel{\mathrm{mL}} \times \underbrace{\frac{1.05\ \mathrm{g}}{1\ \cancel{\mathrm{mL}}}}_{\text{density}} = 15.8\ \text{g of saline solution}$$

Milliliters cancel.

The answer, 15.8 g, is rounded to three significant figures to match the number of significant figures in both factors in the problem.

**Practice Problem 1.21**

Calculate the mass in grams of 10.0 mL of diethyl ether, an anesthetic that has a density of 0.713 g/mL.

**Practice Problem 1.22**

If a household recycles 10.5 kg of plastic bottles composed of PET in a month, and the density of PET is 1.38 g/mL, what volume in liters does this represent?

More Practice: Try Problems 1.48, 1.49.

**HEALTH NOTE**

The specific gravity of a urine sample is measured to check if a patient has an imbalance in metabolism. *Nancy R. Cohen/Photodisc/Getty Images*

## 1.10B Specific Gravity

***Specific gravity*** **is a quantity that compares the density of a substance with the density of water at 4 °C.**

$$\text{specific gravity} = \frac{\text{density of a substance (g/mL)}}{\text{density of water (g/mL)}}$$

Unlike most other quantities, specific gravity is a quantity without units, because the units in the numerator (g/mL) cancel the units in the denominator (g/mL). Because the density of water is 1.00 g/mL at 4 °C, **the specific gravity of a substance equals its density, but it contains no units.** For example, if the density of a liquid is 1.5 g/mL at 4 °C, its specific gravity is 1.5.

The specific gravity of urine samples is often measured in a hospital lab. Normal urine has a density in the range of 1.003–1.030 g/mL (Table 1.6), so it has a specific gravity in the range of 1.003–1.030. Consistently high or low values can indicate an imbalance in metabolism. For example, the specific gravity of urine samples from patients with poorly controlled diabetes is abnormally high, because a large amount of glucose is excreted in the urine.

**Concept Check 1.10**

(a) If the density of a liquid is 0.80 g/mL, what is its specific gravity? (b) If the specific gravity of a substance is 2.3, what is its density?

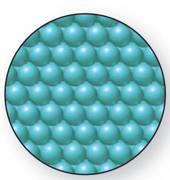
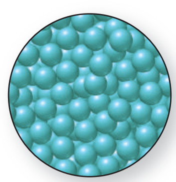
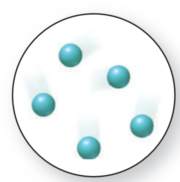
# CHAPTER REVIEW

## KEY TERMS

| | | | |
|---|---|---|---|
| Celsius scale (1.9) | English system of measurement (1.4) | Liter (1.4) | Scientific notation (1.6) |
| Chemical properties (1.2) | Exact number (1.5) | Mass (1.4) | SI units (1.4) |
| Chemistry (1.1) | Fahrenheit scale (1.9) | Matter (1.1) | Significant figures (1.5) |
| Compound (1.3) | Gas (1.2) | Meter (1.4) | Solid (1.2) |
| Conversion factor (1.7) | Gram (1.4) | Metric system (1.4) | Specific gravity (1.10) |
| Cubic centimeter (1.4) | Inexact number (1.5) | Mixture (1.3) | States of matter (1.2) |
| Density (1.10) | Kelvin scale (1.9) | Physical properties (1.2) | Temperature (1.9) |
| Element (1.3) | Liquid (1.2) | Pure substance (1.3) | Weight (1.4) |

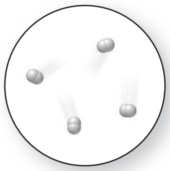
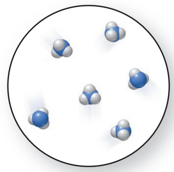
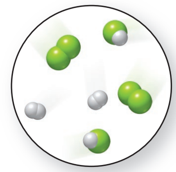
## KEY CONCEPTS

**1 Three states of matter (1.2)**

| 1 **Solid** | 2 **Liquid** | 3 **Gas** |
|---|---|---|
| • The **solid state** is composed of highly organized particles that lie close together. | • The **liquid state** is composed of particles that lie close together but are less organized than the solid state. | • The **gas state** is composed of highly disorganized particles that lie far apart. |
|  |  |  |

**2 Classification of matter (1.3)**

Matter can be classified as a **pure substance** or a **mixture.**

| 1 Pure substance—Element | 2 Pure substance—Compound | 3 Mixture |
|---|---|---|
| • 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 from two or more elements. | • A **mixture** is composed of more than one substance. |
|  |  |  |

**3 Significant figures (1.5)**

Significant figures are all of the digits in a measured number, including one estimated digit.

| 1 **All nonzero digits are significant.** | | 2 **A zero is significant when it occurs between two nonzero digits, or at the end of a number *with* a decimal point.** | | 3 **A zero is *not* significant at the beginning of a number or at the end of a number *without* a decimal point.** | |
|---|---|---|---|---|---|
| 62.34 | 62,443 | 3.02 | 5,280. | 0.05 | 200 |
| four significant figures | five significant figures | three significant figures | four significant figures | one significant figure | one significant figure |

## KEY EQUATIONS

**1 Converting temperature from one scale to another (1.9)**

To convert from Celsius to Fahrenheit:

$$T_F = \frac{9}{5}(T_C) + 32$$

To convert from Fahrenheit to Celsius:

$$T_C = \frac{5}{9}(T_F - 32)$$

To convert from Celsius to Kelvin:

$$T_K = T_C + 273.15$$

To convert from Kelvin to Celsius:

$$T_C = T_K - 273.15$$

**2 Density and specific gravity (1.10)**

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

$$\text{specific gravity} = \frac{\text{density of a substance (g/mL)}}{\text{density of water (g/mL)}}$$

## KEY SKILLS

**1 Rounding off a number (1.5B)**

1 **If a number is to have *two* significant figures, look at the *third* significant number from the left. If the third number is 4 or fewer, drop it and all remaining numbers to the right.**

0.010 023 9 → 0.010

The third number from the left is 4 or fewer.

two significant figures

2 **If the *third* significant number from the left is 5 or greater, round the number up by adding one to the second digit.**

2.7983 → 2.8

The third number from the left is 5 or greater.

two significant figures

**2 Determining significant figures in a calculation**

1 **In multiplication or division, the answer has the same number of significant figures as the original number with the *fewest significant figures.*** (1.5B)

23.2 × 1.1 = 25.52 $\xrightarrow{\text{round off}}$ 26

1.1: two significant figures

25.52: The third number from the left is 5 or greater.

26: two significant figures

2 **In addition and subtraction, the answer has the same number of decimal places as the original number with the *fewest decimal places.*** (1.5C)

3.52 + 2.2 = 5.72 $\xrightarrow{\text{round off}}$ 5.7

2.2: one digit after the decimal point

5.72: last significant figure

5.7: one digit after the decimal point

### 3 Using scientific notation (1.6)

Scientific notation is a method of writing a number as $y \times 10^x$, where $y$ is a number between 1 and 10, and $x$ is a positive or negative exponent.

| 1 To convert a standard number to a number in scientific notation: | 2 To convert a number in scientific notation to a standard number: |
|---|---|
| • Move the decimal point to give a number between 1 and 10.<br>• Multiply the result by $10^x$, where $x$ is the number of places the decimal point was moved.<br>• When the decimal point is moved to the ***left,*** $x$ is **positive.** When the decimal point is moved to the ***right,*** $x$ is **negative.**<br>0.000 032 = 3.2 × $10^{-5}$<br>Move the decimal point **five** places to the **right.** | • Move the decimal point to the ***right*** for **positive** exponents, and to the ***left*** for **negative** exponents.<br>1.400 × $10^{2}$ 1.400 → 140.0<br>Move the decimal point to the **right two** places. |

### 4 Using a conversion factor to solve an equation (1.7, 1.8); example: determining how many liters are contained in 4.0 qt of milk

| 1 Identify the original quantity and the desired quantity. | 2 Write out the conversion factor that places the unwanted unit in the denominator. | 3 Solve the problem so that the unwanted unit cancels. |
|---|---|---|
| 4.0 qt (original quantity) ? L (**desired** quantity) | $\frac{1\ \text{L}}{1.057\ \text{qt}}$ ← **desired** unit / ← unwanted unit<br>**conversion factor** | $4.0\ \cancel{\text{qt}} \times \frac{1\ \text{L}}{1.057\ \cancel{\text{qt}}} = 3.8\ \text{L}$<br>**conversion factor**<br>Quarts cancel. |

### 5 Converting temperature from one scale to another (1.9); example: converting 92 °F to $T_C$ and $T_K$

| 1 Convert the Fahrenheit temperature to degrees Celsius using the equation $T_C = 5/9(T_F - 32)$. | 2 Convert the Celsius temperature to kelvins by adding 273.15. |
|---|---|
| $T_C = \frac{5}{9}(T_F - 32) = \frac{5}{9}(92 - 32) = 33\ °\text{C}$ | $T_K = T_C + 273.15 = 33 + 273.15 = 306\ \text{K}$ |

### 6 Using density to solve a problem (1.10A); example: determining the mass of 20.0 mL of an antibiotic solution (density = 1.21 g/mL)

| 1 Identify the original quantity and the desired quantity. | 2 Use the density as a conversion factor to place the unwanted unit (mL) in the denominator. | 3 Solve the problem so that the unwanted unit (mL) cancels. |
|---|---|---|
| 20.0 mL (original quantity) ? g (**desired** quantity) | $\frac{1.21\ \text{g}}{1\ \text{mL}}$ ← **desired** unit / ← unwanted unit<br>**conversion factor** | $20.0\ \cancel{\text{mL}} \times \frac{1.21\ \text{g}}{1\ \cancel{\text{mL}}} = 24.2\ \text{g solution}$<br>**conversion factor**<br>mL cancel. |

## CHAPTER 1 SELF-TEST

The Self-Test consists of short-answer questions that test an understanding of definitions, equations, and other material encountered within the chapter. Answers to each question are provided at the end of the chapter. For any questions you are unable to answer correctly, revisit the appropriate section(s) of the text; then proceed to working the end-of-chapter problems to further reinforce your understanding of the chapter material.

**1.** Which term from "Key Terms" answers each question?
   a. Which state of matter consists of close, randomly moving particles that takes on the shape of the container it occupies?
   b. What physical property relates the mass of an object to its volume?
   c. What type of pure substance cannot be broken down into simpler substances by a chemical reaction?
   d. On what temperature scale does water freeze at 0° and boil at 100°?

**2.** How many significant figures does each number contain: (a) 54,000; (b) 0.067 00; (c) 1,109.0?

**3.** What term is used for each unit: (a) a thousand seconds; (b) a hundredth of a meter?

**4.** Which number represents 10,400 written in scientific notation: (a) $1.04 \times 10^{-4}$; (b) $1.04 \times 10^4$; (c) $1.0400 \times 10^4$; (d) $10.4 \times 10^3$?

**5.** Round off each number to two significant figures: (a) 2.7983; (b) 0.010 023 9; (c) 42,980.0

**6.** The density of liquid **A** is half the density of liquid **B.** What volume of liquid **B** has the same mass as 50 mL of liquid **A?**

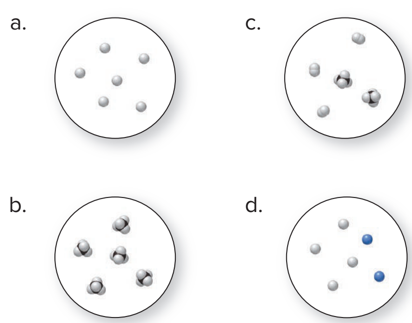
**7.** What is the product of the following calculation, written with the proper number of significant figures; 4.29 × 0.023 =: (a) 0.098 67; (b) 0.0987; (c) 0.099?

**8.** To determine how many liters are contained in 1.0 gal of milk, fill in quantities (a) and (b) needed to solve the following equation.

$$1.0 \text{ gal} \times \frac{4 \text{ qt}}{(a)} \times \frac{1 \text{ L}}{(b)} = 3.8 \text{ L}$$

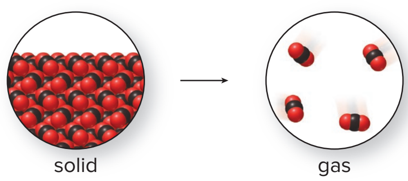
## UNDERSTANDING KEY CONCEPTS

**1.1** Classify each example of molecular art as a pure element, a pure compound, or a mixture.

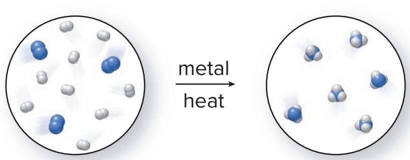



**1.2** (a) Which representation(s) in Problem 1.1 illustrate a mixture of two elements? (b) Which representation(s) in Problem 1.1 illustrate a mixture of a compound and an element?

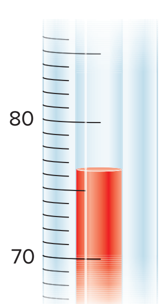
**1.3** When a chunk of dry ice (solid carbon dioxide) is placed out in the air, the solid gradually disappears and a gas is formed above the solid. Does the molecular art drawn below indicate that a chemical or physical change has occurred? Explain your choice.




**1.4** The inexpensive preparation of nitrogen-containing fertilizers begins with mixing together two elements, hydrogen and nitrogen, at high temperature and pressure in the presence of a metal. Does the molecular art depicted below indicate that a chemical or physical change occurs under these conditions? Explain your choice.




**1.5**

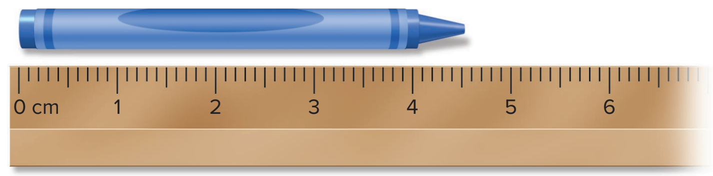



a. What is the temperature on the given Fahrenheit thermometer?

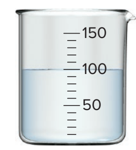
b. How many significant figures does your answer contain?

c. Convert this temperature into $T_C$.

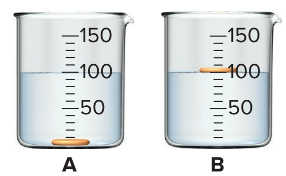
**1.6** (a) What is the length of the given crayon in centimeters? (b) How many significant figures does this value contain? (c) Convert this value to meters, and write the answer in scientific notation.




**1.7** The given beaker contains 100 mL of water. Draw an illustration for what would be observed in each circumstance. (a) Hexane (50 mL, density = 0.65 g/mL) is added. (b) Dichloromethane (50 mL, density = 1.33 g/mL) is added.




**1.8** (a) What can be said about the density of the liquid in beaker **A,** if the object in the beaker has a density of 2.0 g/cc? (b) What can be said about the density of the liquid in beaker **B,** if the object has a density of 0.90 g/cc?




**1.9** A blood vessel is 0.40 µm in diameter. (a) Convert this quantity to meters and write the answer in scientific notation. (b) Convert this quantity to inches and write the answer in scientific notation.

**1.10** On admission to the hospital, a patient weighed 60.6 kg and was 67 in. tall. (a) What is the weight of the patient in pounds? (b) What is the height of the patient in centimeters?

**1.11** What is the height in centimeters of a child who is 50. in. tall?

**1.12** If the human body has 5.0 qt of blood that contains $2 \times 10^{13}$ red blood cells, how many red blood cells are present in each µL of blood?

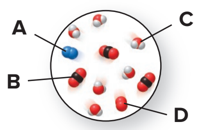
**1.13** A woman was told to take a dose of 1.5 g of calcium daily. How many 500-mg tablets should she take?

**1.14** The recommended daily calcium intake for a woman over 50 years of age is 1,200 mg. If one cup of milk has 306 mg of calcium, how many cups of milk provide this amount of calcium? (b) How many milliliters of milk does this correspond to? (1 cup = 8 fl oz)

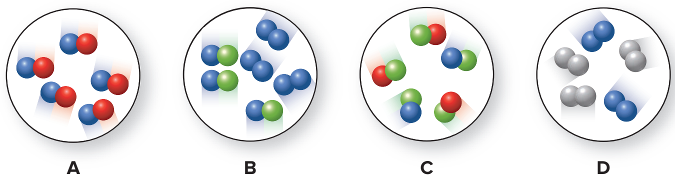
# ADDITIONAL PROBLEMS

## Matter

**1.15** Label each component in the molecular art as an element or a compound.




**1.16** (a) Which representation(s) illustrate a mixture of two elements? (b) Which representation(s) illustrate a mixture of a compound and an element?




**1.17** Classify each process as a chemical or physical change.
- a. dissolving calcium chloride in water
- b. burning gasoline to power a car
- c. heating wax so that it melts

**1.18** Classify each process as a chemical or physical change.
- a. the condensation of water on the outside of a cold glass
- b. mixing a teaspoon of instant coffee with hot water
- c. baking a cake

## Measurement

**1.19** Which quantity in each pair is larger?
- a. 5 mL or 5 dL
- b. 10 mg or 10 μg
- c. 5 cm or 5 mm
- d. 10 Ms or 10 ms

**1.20** Which quantity in each pair is larger?
- a. 10 km or 10 m
- b. 10 L or 10 mL
- c. 10 g or 10 μg
- d. 10 cm or 10 mm

**1.21** Label each quantity as an exact or inexact number.
- a. A recipe requires 10 cloves of garlic and two tablespoons of oil.
- b. A child fell and had a 4-cm laceration that required 12 stitches.

**1.22** Rank the quantities in each group from smallest to largest.
- a. 100 μL, 100 dL, and 100 mL
- b. 10 g, 100 mg, and 0.1 kg
- c. 1 km, 100 m, and 1,000 cm

## Significant Figures

**1.23** How many significant figures does each number contain?
- a. 16.00
- b. 160
- c. 0.001 60
- d. 1,600,000
- e. 0.1600
- f. $1.060 \times 10^{10}$

**1.24** How many significant figures does each number contain?
- a. 160.
- b. 160.0
- c. 0.000 16
- d. 1,600.
- e. 1.060
- f. $1.600 \times 10^{-10}$

**1.25** Round each number to three significant figures.
- a. 25,401
- b. 1,248,486
- c. 0.001 265 982
- d. 0.123 456

**1.26** Round each number in Problem 1.25 to four significant figures.

**1.27** Carry out each calculation and report the answer using the proper number of significant figures.
- a. $53.6 \times 0.41$
- b. $25.825 - 3.86$
- c. $65.2 \div 12$
- d. $41.0 + 9.135$

**1.28** Carry out each calculation and report the answer using the proper number of significant figures.
- a. $49{,}682 \times 0.80$
- b. $66.815 + 2.82$
- c. $1{,}000 \div 2.34$
- d. $21 - 0.88$

## Scientific Notation

**1.29** Write each quantity in scientific notation.
- a. 1,234 g
- b. 0.000 016 2 m
- c. 5,244,000 L
- d. 0.005 62 g

**1.30** Write each quantity in scientific notation.
- a. 0.001 25 m
- b. 8,100,000,000 lb
- c. 54,235.6 m
- d. 0.000 001 899 L

**1.31** Convert each number to its standard form.
- a. $3.4 \times 10^{8}$
- b. $5.822 \times 10^{-5}$
- c. $3 \times 10^{2}$
- d. $6.86 \times 10^{-8}$

**1.32** Convert each number to its standard form.
- a. $4.02 \times 10^{10}$
- b. $2.46 \times 10^{-3}$
- c. $6.86 \times 10^{9}$
- d. $1.00 \times 10^{-7}$

**1.33** Which number in each pair is larger?
- a. $4.44 \times 10^{3}$ or $4.8 \times 10^{2}$
- b. $5.6 \times 10^{-6}$ or $5.6 \times 10^{-5}$
- c. $1.3 \times 10^{8}$ or 52,300,000
- d. $9.8 \times 10^{-4}$ or 0.000 089

**1.34** Rank the numbers in each group from smallest to largest.
- a. $5.06 \times 10^{6}$, $7 \times 10^{4}$, and $2.5 \times 10^{8}$
- b. $6.3 \times 10^{-2}$, $2.5 \times 10^{-4}$, and $8.6 \times 10^{-6}$

**1.35** Write the recommended daily intake of each nutrient in scientific notation.
- a. 0.000 400 g of folate
- b. 0.002 g of copper
- c. 0.000 080 g of vitamin K
- d. 3,400 mg of chloride

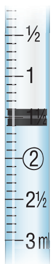
**1.36** A picosecond is one trillionth of a second (0.000 000 000 001 s). (a) Write this number in scientific notation. (b) How many picoseconds are there in one second? Write this answer in scientific notation.

## Problem Solving Using Conversion Factors

**1.37** Carry out each of the following conversions.
- a. 300 g to mg
- b. 2 L to μL
- c. 5.0 cm to m
- d. 2 ft to m

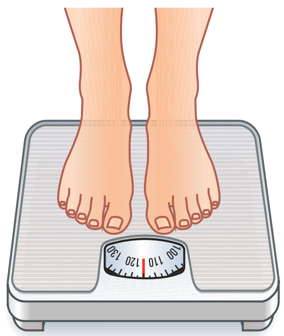
**1.38** Carry out each of the following conversions.

a. 25 μL to mL
b. 35 kg to g
c. 300 mL to qt
d. 3 cups to L

**1.39** 


a. What is the volume of liquid contained in the given 3-mL syringe?
b. Convert this value to liters and write the answer in scientific notation.

**1.40** What is the mass in kilograms of an individual whose weight in pounds is shown on the given scale?



**1.41** Carry out each of the following conversions.

a. What is the mass in kilograms of an individual who weighs 234 lb?
b. A patient required 3.0 pt of blood during surgery. How many liters does this correspond to?
c. A patient had a body temperature of 37.7 °C. What is his body temperature in $T_F$?

**1.42** Carry out each of the following conversions.

a. What is the mass in pounds of an individual who weighs 53.2 kg?
b. What is the height in inches of a child who is 90. cm tall?
c. A patient had a body temperature of 103.5 °F. What is his body temperature in $T_C$?

**1.43** The concentration of mercury, a toxic pollutant, in some yellowfin tuna is 354 mg in $1.0 \times 10^6$ g of fish. The Food and Drug Administration recommends that women who are pregnant should eat no more than 6.0 oz of yellowfin tuna per week. How much mercury (in milligrams) would an individual ingest in one serving of fish?

*Nishihama/Shutterstock*

**1.44** Honeydew melons are high in potassium, a nutrient needed for proper fluid balance [875 mg in one serving (one-quarter of a medium melon)]. If an individual eats one serving each day for a week, how many grams of potassium are obtained?

*Brent Hofacker/Shutterstock*

## Temperature

**1.45** Carry out each of the following temperature conversions.

a. An over-the-counter pain reliever melts at 53 °C. Convert this temperature to $T_F$ and $T_K$.
b. A cake is baked at 350. °F. Convert this temperature to $T_C$ and $T_K$.

**1.46** Which temperature in each pair is higher?

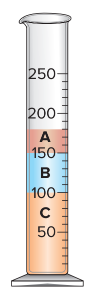
a. –10 °C or 10 °F
b. –50 °C or –50 °F

## Density and Specific Gravity

**1.47** If a urine sample has a mass of 122 g and a volume of 121 mL, what is its density in g/mL?

**1.48** The density of sucrose, table sugar, is 1.56 g/cc. What volume (in cubic centimeters) does 20.0 g of sucrose occupy?

**1.49** Isooctane is a high-octane component of gasoline. If the density of isooctane is 0.692 g/mL, what is the mass of 220 mL of gasoline?

**1.50** 


A graduated cylinder contains three liquids: water (density = 1.00 g/mL), corn syrup (density = 1.37 g/mL), and corn oil (density = 0.93 g/mL). (a) Identify the liquid that corresponds to **A, B,** and **C** in the graduated cylinder. (b) How many grams of each liquid are present?

**1.51** Which is the upper layer when each of the following liquids is added to water?

a. olive oil (density = 0.92 g/mL)
b. chloroform (density = 1.49 g/mL)

**1.52** (a) What is the specific gravity of mercury, the liquid used in thermometers, if it has a density of 13.6 g/mL? (b) What is the density of ethanol if it has a specific gravity of 0.789?