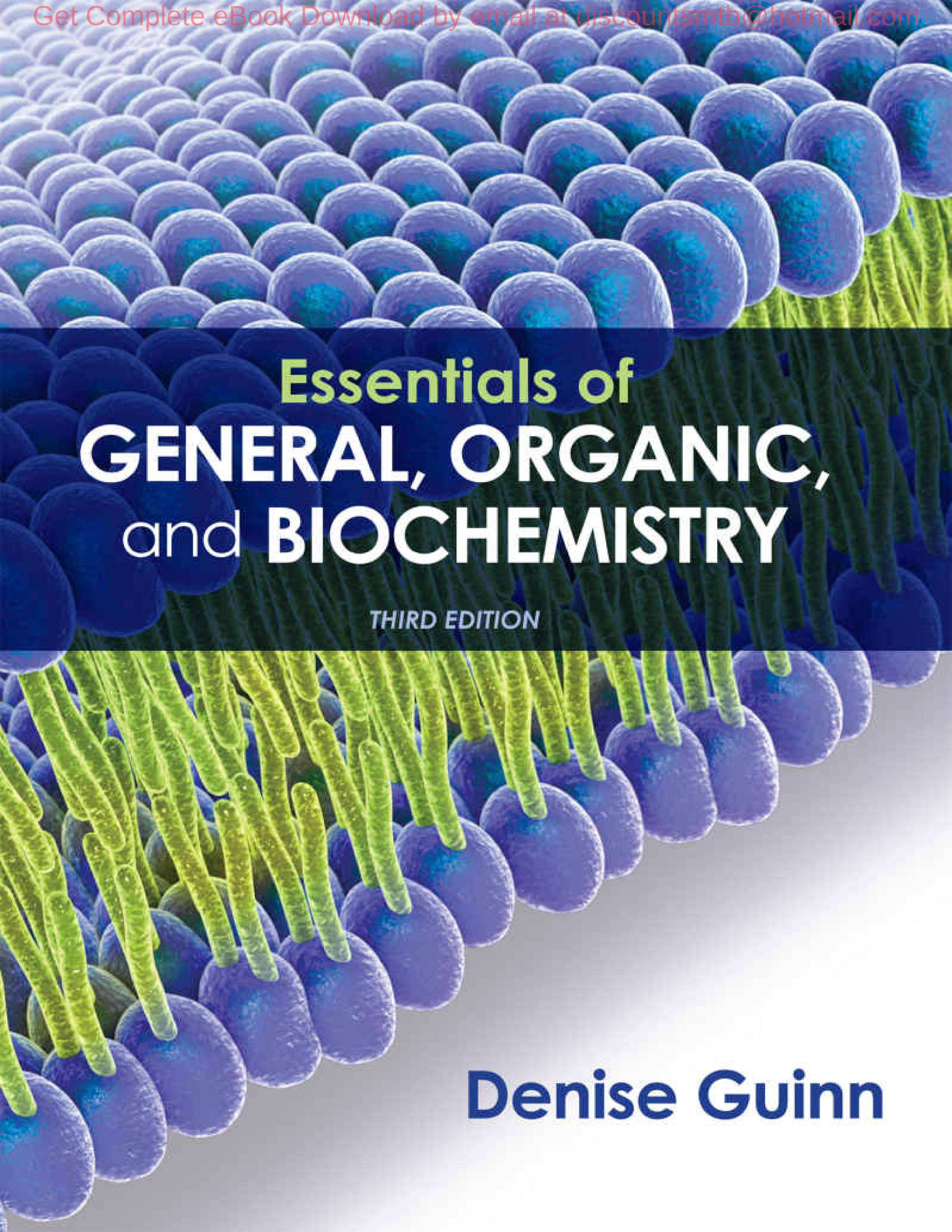
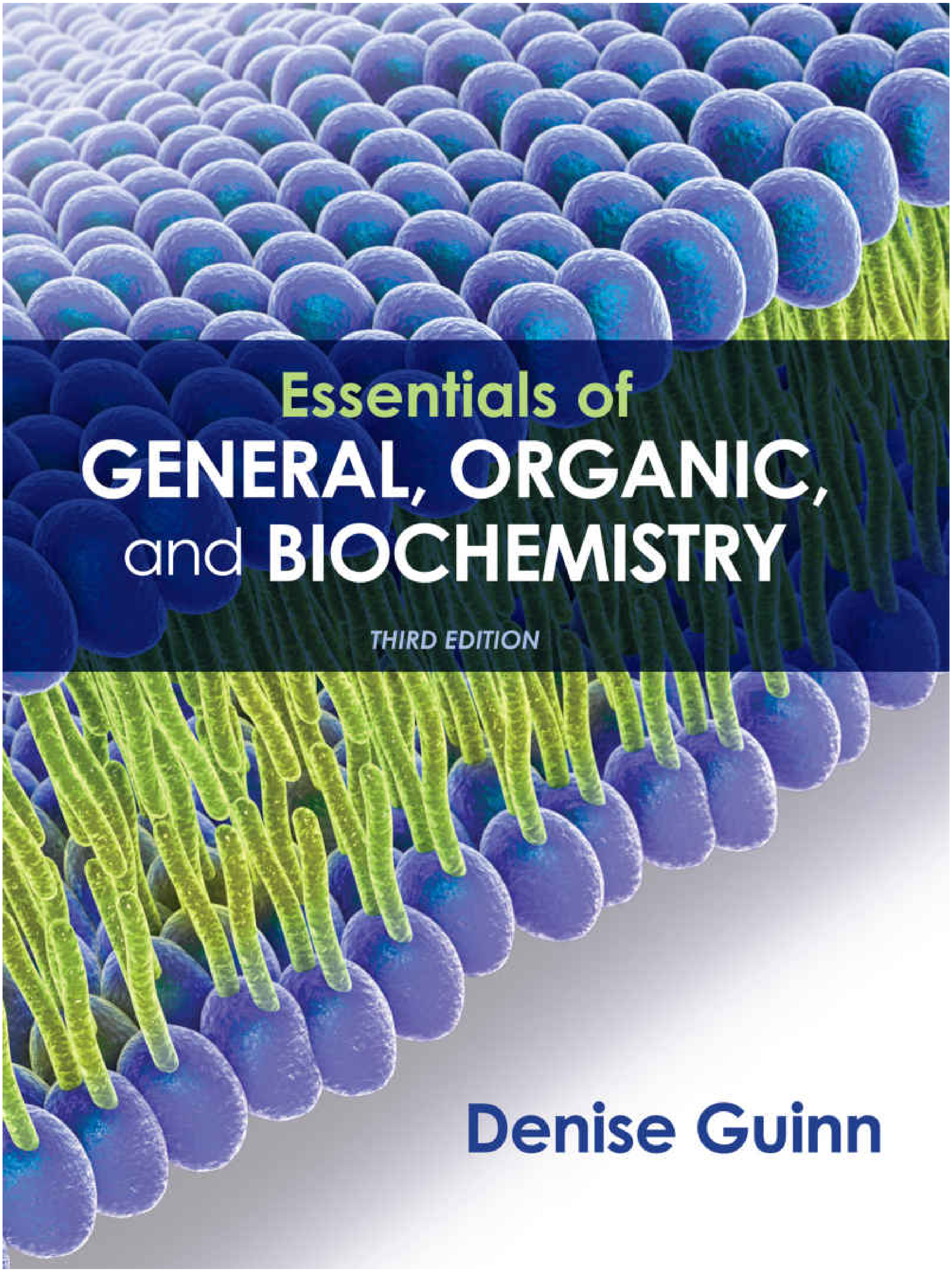


Essentials of GENERAL, ORGANIC, and BIOCHEMISTRY

THIRD EDITION

Denise Guinn


Essentials of
GENERAL, ORGANIC,
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Periodic Table of the Elements

Key: Atomic number 6; Atomic Symbol C; Atomic mass* 12.01; Name Carbon

Main group elements (1A, 2A; 3A–8A) — Transition metals (3B–2B)

Periods	1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1	1 H 1.008 Hydrogen																	2 He 4.003 Helium
2	3 Li 6.941 Lithium	4 Be 9.012 Beryllium											5 B 10.81 Boron	6 C 12.01 Carbon	7 N 14.01 Nitrogen	8 O 16.00 Oxygen	9 F 19.00 Fluorine	10 Ne 20.18 Neon
3	11 Na 22.99 Sodium	12 Mg 24.31 Magnesium											13 Al 26.98 Aluminum	14 Si 28.09 Silicon	15 P 30.97 Phosphorus	16 S 32.07 Sulfur	17 Cl 35.45 Chlorine	18 Ar 39.95 Argon
4	19 K 39.10 Potassium	20 Ca 40.08 Calcium	21 Sc 44.96 Scandium	22 Ti 47.87 Titanium	23 V 50.94 Vanadium	24 Cr 52.00 Chromium	25 Mn 54.94 Manganese	26 Fe 55.85 Iron	27 Co 58.93 Cobalt	28 Ni 58.69 Nickel	29 Cu 63.55 Copper	30 Zn 65.41 Zinc	31 Ga 69.72 Gallium	32 Ge 72.61 Germanium	33 As 74.92 Arsenic	34 Se 78.96 Selenium	35 Br 79.90 Bromine	36 Kr 83.80 Krypton
5	37 Rb 85.47 Rubidium	38 Sr 87.62 Strontium	39 Y 88.91 Yttrium	40 Zr 91.22 Zirconium	41 Nb 92.91 Niobium	42 Mo 95.94 Molybdenum	43 Tc (98) Technetium	44 Ru 101.1 Ruthenium	45 Rh 102.9 Rhodium	46 Pd 106.4 Palladium	47 Ag 107.9 Silver	48 Cd 112.4 Cadmium	49 In 114.8 Indium	50 Sn 118.7 Tin	51 Sb 121.8 Antimony	52 Te 127.6 Tellurium	53 I 126.9 Iodine	54 Xe 131.3 Xenon
6	55 Cs 132.9 Cesium	56 Ba 137.3 Barium	57 La† 138.9 Lanthanum	72 Hf 178.5 Hafnium	73 Ta 180.9 Tantalum	74 W 183.9 Tungsten	75 Re 186.2 Rhenium	76 Os 190.2 Osmium	77 Ir 192.2 Iridium	78 Pt 195.1 Platinum	79 Au 197.0 Gold	80 Hg 200.6 Mercury	81 Tl 204.4 Thallium	82 Pb 207.2 Lead	83 Bi 209.0 Bismuth	84 Po (209) Polonium	85 At (210) Astatine	86 Rn (222) Radon
7	87 Fr (223) Francium	88 Ra (226) Radium	89 Ac‡ (227) Actinium	104 Rf (261) Rutherfordium	105 Db (262) Dubnium	106 Sg (266) Seaborgium	107 Bh (267) Bohrium	108 Hs (277) Hassium	109 Mt (268) Meitnerium	110 Ds (269) Darmstadtium	111 Rg (272) Roentgenium	112 Cn (285) Copernicium	113 Nh (284) Nihonium	114 Fl (289) Flerovium	115 Mc (288) Moscovium	116 Lv (293) Livermorium	117 Ts (294) Tennessine	118 Og (294) Oganesson

† Lanthanides	58 Ce 140.1 Cerium	59 Pr 140.9 Praseodymium	60 Nd 144.2 Neodymium	61 Pm (145) Promethium	62 Sm 150.4 Samarium	63 Eu 152.0 Europium	64 Gd 157.3 Gadolinium	65 Tb 158.9 Terbium	66 Dy 162.5 Dysprosium	67 Ho 164.9 Holmium	68 Er 167.3 Erbium	69 Tm 168.9 Thulium	70 Yb 173.0 Ytterbium	71 Lu 175.0 Lutetium
‡ Actinides	90 Th 232.0 Thorium	91 Pa 231.0 Protactinium	92 U 238.0 Uranium	93 Np 237.0 Neptunium	94 Pu (244) Plutonium	95 Am (243) Americium	96 Cm (247) Curium	97 Bk (247) Berkelium	98 Cf (251) Californium	99 Es (252) Einsteinium	100 Fm (257) Fermium	101 Md (258) Mendelevium	102 No (259) Nobelium	103 Lr (260) Lawrencium

Metals
Metalloids
Nonmetals

* Atomic masses given in parentheses indicate the atomic mass of the most stable isotope.

The Elements

Element	Symbol	Atomic number	Atomic mass (amu)
actinium	Ac	89	(227)
aluminum	Al	13	26.98
americium	Am	95	(243)
antimony	Sb	51	121.8
argon	Ar	18	39.95
arsenic	As	33	74.92
astatine	At	85	(210)
barium	Ba	56	137.3
berkelium	Bk	97	(247)
beryllium	Be	4	9.012
bismuth	Bi	83	209.0
bohrium	Bh	107	(267)
boron	B	5	10.81
bromine	Br	35	79.90
cadmium	Cd	48	112.4
calcium	Ca	20	40.08
californium	Cf	98	(251)
carbon	C	6	12.01
cerium	Ce	58	140.1
cesium	Cs	55	132.9
chlorine	Cl	17	35.45
chromium	Cr	24	52.00
cobalt	Co	27	58.93
copernicium	Cn	112	285
copper	Cu	29	63.55
curium	Cm	96	(247)
darmstadtium	Ds	110	(269)
dubnium	Db	105	(262)
dysprosium	Dy	66	162.5
einsteinium	Es	99	(252)
erbium	Er	68	167.3
europium	Eu	63	152.0
fermium	Fm	100	(257)
flerovium	Fl	114	289
fluorine	F	9	19.00
francium	Fr	87	(223)
gadolinium	Gd	64	157.3
gallium	Ga	31	69.72

germanium	Ge	32	72.40
gold	Au	79	197.0
hafnium	Hf	72	178.5
hassium	Hs	108	(277)
helium	He	2	4.003
holmium	Ho	67	164.9
hydrogen	H	1	1.008
indium	In	49	114.8
iodine	I	53	126.9
iridium	Ir	77	192.2
iron	Fe	26	55.85
krypton	Kr	36	83.80
lanthanum	La	57	138.9
lawrencium	Lr	103	(260)
lead	Pb	82	207.2
lithium	Li	3	6.941
livermorium	Lv	116	293
lutetium	Lu	71	175.0
magnesium	Mg	12	24.31
manganese	Mn	25	54.94
meitnerium	Mt	109	(268)
mendelevium	Md	101	(258)
mercury	Hg	80	200.6
molybdenum	Mo	42	95.94
moscovium	Mc	115	288
neodymium	Nd	60	144.2
neon	Ne	10	20.18
neptunium	Np	93	237.0
nickel	Ni	28	58.69
nihonium	Nh	113	284
niobium	Nb	41	92.91
nitrogen	N	7	14.01
nobelium	No	102	(259)
oganesson	Og	118	294
osmium	Os	76	190.2
oxygen	O	8	16.00
palladium	Pd	46	106.4
phosphorus	P	15	30.97
platinum	Pt	78	195.1
plutonium	Pu	94	(244)
polonium	Po	84	(209)

potassium	K	19	39.10
praseodymium	Pr	59	140.9
promethium	Pm	61	(145)
protactinium	Pa	91	231.0
radium	Ra	88	(226)
radon	Rn	86	(222)
rhenium	Re	75	186.2
rhodium	Rh	45	102.9
roentgenium	Rg	111	(272)
rubidium	Rb	37	85.47
ruthenium	Ru	44	101.1
rutherfordium	Rf	104	(261)
samarium	Sm	62	150.4
scandium	Sc	21	44.96
seaborgium	Sg	106	(266)
selenium	Se	34	78.96
silicon	Si	14	28.09
silver	Ag	47	107.9
sodium	Na	11	22.99
strontium	Sr	38	87.62
sulfur	S	16	32.07
tantalum	Ta	73	180.9
technetium	Tc	43	(98)
tellurium	Te	52	127.6
tennessine	Ts	117	294
terbium	Tb	65	158.9
thallium	Tl	81	204.4
thorium	Th	90	232.0
thulium	Tm	69	168.9
tin	Sn	50	118.7
titanium	Ti	22	47.87
tungsten	W	74	183.9
uranium	U	92	238.0
vanadium	V	23	50.94
xenon	Xe	54	131.3
ytterbium	Yb	70	173.0
yttrium	Y	39	88.91
zinc	Zn	30	65.41
zirconium	Zr	40	91.22

Note: Parentheses ( ) denote the most stable isotope of a radioactive element.

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The College of New Rochelle

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PRINTING AND BINDING: LSC Communications INTERIOR ICONS: *Core Concept*: bearsky23/Shutterstock; *Math Tip*: siridhata/Shutterstock COVER AND TITLE PAGE ILLUSTRATION: Echo Medical Media

Library of Congress Control Number: 2018959141

ISBN: 978-1-319-26756-8 (mobi)

1 2 3 4 5 6 23 22 21 20 19 18

W. H. Freeman and Company

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To my sons, Charlie and Scott, for their unwavering support.

About the Author

Chris Carroll Photography

Denise Guinn received her B.A. in chemistry from the University of California at San Diego and her Ph.D. in organic chemistry from the University of Texas at Austin. She was a National Institutes of Health postdoctoral fellow at Harvard University before joining Abbott Laboratories as a Research Scientist in the Pharmaceutical Products Discovery Group. In 1992, Dr. Guinn joined the faculty at Regis University in Denver, Colorado, as Clare Boothe Luce Assistant Professor of Chemistry, where she taught courses in general chemistry, organic chemistry, and the general, organic, and biochemistry course for nursing and allied health majors. In 2008, she joined the chemistry department at The College of New Rochelle, in New York where she teaches organic chemistry, biochemistry, and the one-semester GOB course for nursing students. She has published in the *Journal of Organic Chemistry,* the *Journal of the American Chemical Society,* and the *Journal of Medicinal Chemistry.* She currently resides in Nyack, New York with her dog Puck.

Brief Contents

Contents

A letter from the author

In teaching the general, organic, and biochemistry course for the past 26 years, it has been a great pleasure to be part of this course's evolution. When we wrote the first edition of *Essentials of General, Organic, and Biochemistry,* our goal was to make it obvious to students why chemistry is a cornerstone in the education of today's health care professionals, by using health and medicine as the framework for learning the fundamentals of chemistry. The third edition contains updated and additional medical applications integrated into the material, and is now organized into smaller chapters. We have heard from instructors that integrating medical applications throughout the text effectively engages students in the course early on, motivating them to learn the fundamental concepts of chemistry. Based on feedback from hundreds of instructors who teach this course, the chapters on solutions and acid-base chemistry have been rewritten and now appear before the chapters on organic chemistry and the four chapters on organic chemistry are in sequence. The third edition also contains additional in-chapter worked exercises and practice exercises to help students learn problem solving and critical thinking skills. I hope that you feel, as I do, that the third edition has retained the elements of the previous editions that worked well, while incorporating some organizational changes that better support student learning.

Preface

The Essential Chemistry for Health Careers and Everyday Life

Guinn's *Essentials of General, Organic, and Biochemistry* uses health and medicine as the framework for learning the fundamentals of chemistry. The newly revised third edition focuses on core concepts and necessary math skills with a revamped organization. Easily digestible content is served in shorter, concise chapters, while medical applications make chemistry meaningful for students preparing for future careers in nursing and other allied health professions. Using SaplingPlus and its embedded e-book, students will be able to focus their study with adaptive quizzing and understand the relevance of chemistry through videos, animations, and case studies.

Integrates Health and Medicine

Chemistry is the central science, yet students often struggle to see its connection to their lives and career goals. Examples from medical and allied health fields and other consumer-based examples illustrate the fundamental concepts of chemistry throughout the entire GOB sequence of topics. Content is tailored to motivate students and help them understand how chemistry applies to their lives and majors.

Supports Student Success

The third edition of *Essentials of General, Organic, and Biochemistry* ensures students successfully learn chemistry by reinforcing core concepts, math skills, and problem-solving techniques. Bolstering students' abilities in these areas builds confidence and provides the necessary foundation they need to succeed throughout the course.

Promotes Independent Practice with SaplingPlus

Guinn's third edition of *Essentials of General, Organic, and Biochemistry* is now paired with SaplingPlus, promoting student study and practice. Adaptive LearningCurve assignments help direct purposeful reading and study, while tutorials and case studies help students synthesize and apply their understanding. In addition, students and instructors will have access to the acclaimed Sapling Learning online homework and access to industry-leading peer-to-peer support for help with implementation and technical support.

Integrates Health and Medicine

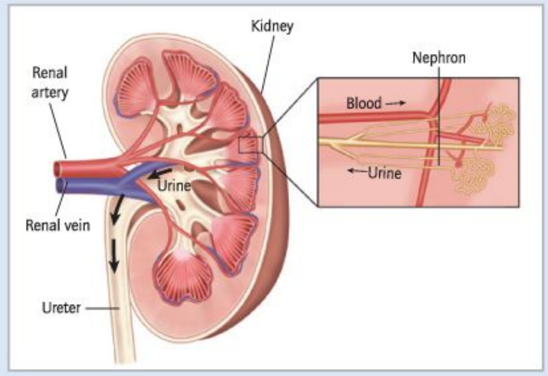
Chemistry is the central science, yet students often struggle to see its connection to their lives and career goals. Examples from medical and allied health fields and other consumer-based examples illustrate the fundamental concepts of chemistry throughout the GOB sequence of topics. Content is tailored to motivate students and help them understand how chemistry applies to their lives and majors.

Concepts in Context: Each chapter opens with a short real-life example of a topic connected to the chapter concepts. These stories immediately immerse the student in a high-interest topic which can be understood through chemistry.

CONCEPTS IN CONTEXT: How the Kidneys Filter Our Blood

Chronic Kidney Disease affects 14 percent of the United States population, and causes more deaths than breast cancer or prostate cancer according to the U.S. Department of Health and Human Services. When a person's kidneys are destroyed by disease, a kidney transplant becomes the only option for the patient's long-term survival. In the interim, the patient must undergo dialysis three or four times a week. Dialysis is a procedure that performs the function of the kidneys. Today, about half a million Americans are on dialysis and 95,000 are actively awaiting a donor kidney according to UNOS (United Network for Organ Sharing).

To understand how the kidneys filter our blood, we need to understand mixtures, the subject of this chapter. Mixtures are a type of matter composed of two or more elements or compounds. Blood is a complex mixture that contains water, ions, elements, and molecules in addition to red blood cells, white



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Authentic Images: Photos related to clinical practice, as well as consumer and health care products, reinforce chemical concepts and show their applications.

Scott Linnett/ZUMA Press/Newscom

8-22 IV solutions of heparin are available as 25,000 units/500 mL of solution, as shown at left. An order is given to infuse 50. units of heparin per hour. At what flow rate, in milliliters per hour, should the IV solution be infused into the patient? Protein concentrations are often reported in international units, IU, or simply "units." Treat international units just like you would the mass of a solute.

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Chemistry in Medicine: Guinn concludes each chapter with an in-depth look at how one or more chemical principles described in the chapter directly applies to a specific health care issue.

Chemistry in Medicine Hemodialysis—Performing the Function of the Kidneys

The kidneys serve the important function of filtering our blood, as described in the *Concepts in Context* at the beginning of the chapter. Kidney failure, known as renal failure, can occur suddenly, referred to as *acute*, or it may be of gradual onset (*chronic*). Acute renal failure can be caused by severe shock, dehydration, a heart attack, or a severe kidney infection. Chronic renal failure can be caused by diabetes, high blood pressure, or certain hereditary factors.

A person with less than 15% kidney function must undergo hemodialysis, a life-support treatment that performs the function of the kidneys, until a donor kidney can be found. Hemodialysis removes waste products and excess water from the blood, while retaining colloids such as proteins and larger particles such as blood cells.

Hemodialysis is typically a 5-hour treatment performed three times a week at a medical center. During the procedure, blood is pumped out of a vein in the forearm and into the *dialyzer*, which filters the blood much like a kidney. The filtered blood is then returned to the body through another vein, as illustrated in Figure 8-23.

The dialyzer contains an aqueous solution, known as the *dialysate*, that is similar in composition to filtered blood serum but hypotonic (less concentrated)

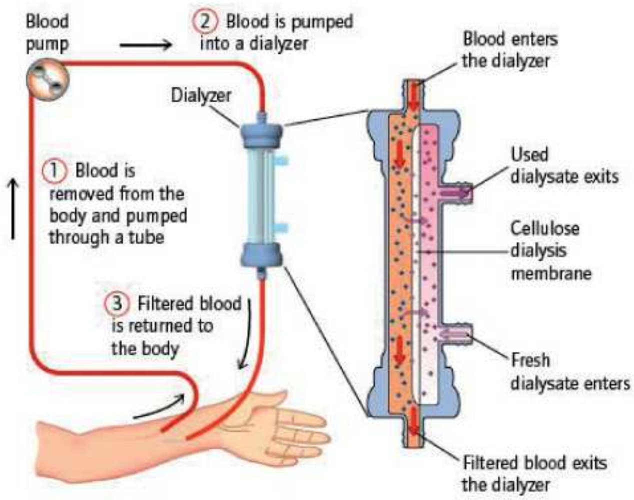

Figure 8-23 Hemodialysis. Blood is removed through a vein and pumped through a cellulose tube in the dialyzer, where it is filtered by dialysis with a solution of dialysate, and then returned to the body through another vein.

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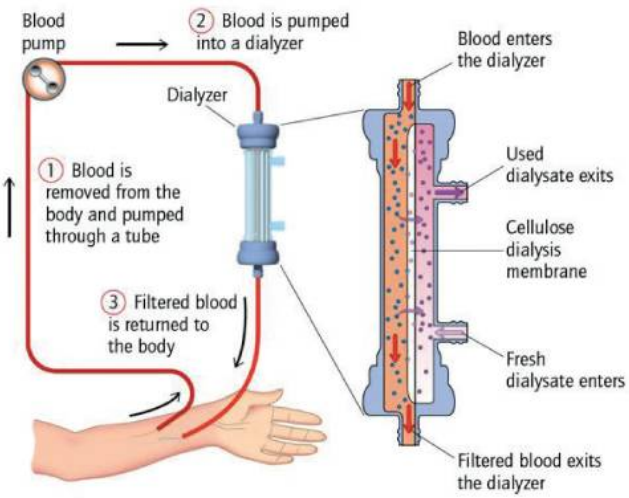
Case Studies: Available as student handouts, case studies allow students to apply what they learn to more challenging problems, connect the chemistry they are learning to medical topics, and provide instructors with tools for active learning. Students also have the opportunity to synthesize their understanding by working through paired case studies in SaplingPlus.

Case Study: Dialysis

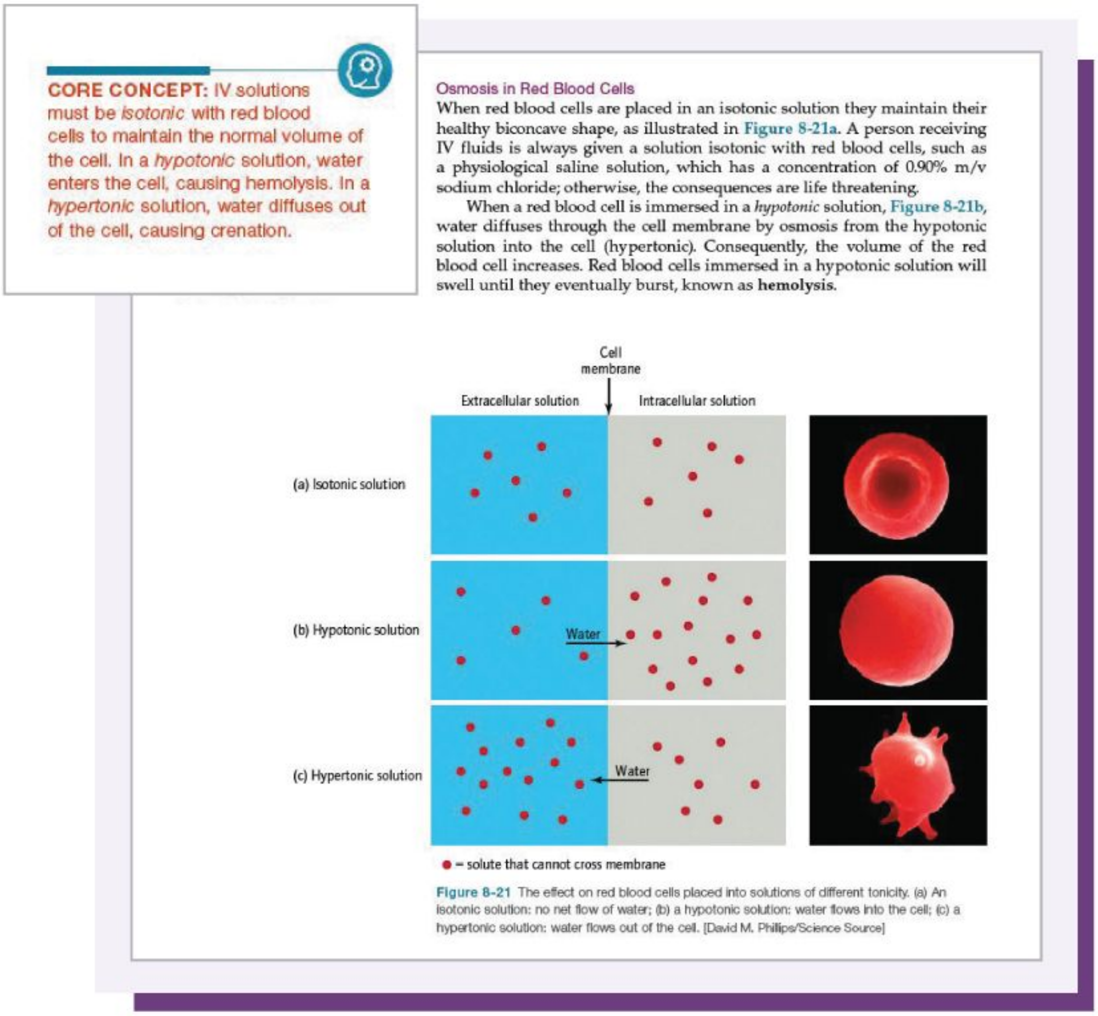
Background: Kidney disease affects more than 31 million people in the United States and 2.6 million people in Canada, and causes more deaths than does breast cancer or prostate cancer. When a person's kidneys are destroyed by disease, a kidney transplant becomes the only option for their long-term survival. In the interim, these people must undergo dialysis, typically three or four times a week. Dialysis is a procedure that performs the function of the kidneys - the removal of waste and water from the body, and maintaining electrolyte balance - performed by a dialyzer (see Figure). Today, more than half a million Americans are on dialysis and more than 100,000 are actively awaiting a donor kidney.



Supports Student Success

The third edition of *Essentials of General, Organic, and Biochemistry* supports students in successfully learning chemistry by reinforcing core concepts, math skills, and problem-solving techniques. Bolstering students' abilities in these areas builds confidence and provides the necessary foundation they need throughout the course.

Core Concepts: These margin notes emphasize important concepts within each chapter and provide a quick study tool for students reviewing chapter content.



CORE CONCEPT: IV solutions must be *isotonic* with red blood cells to maintain the normal volume of the cell. In a *hypotonic* solution, water enters the cell, causing hemolysis. In a *hypertonic* solution, water diffuses out of the cell, causing crenation.

Osmosis in Red Blood Cells

When red blood cells are placed in an isotonic solution they maintain their healthy biconcave shape, as illustrated in Figure 8-21a. A person receiving IV fluids is always given a solution isotonic with red blood cells, such as a physiological saline solution, which has a concentration of 0.90% m/v sodium chloride; otherwise, the consequences are life threatening.

When a red blood cell is immersed in a *hypotonic* solution, Figure 8-21b, water diffuses through the cell membrane by osmosis from the hypotonic solution into the cell (hypertonic). Consequently, the volume of the red blood cell increases. Red blood cells immersed in a hypotonic solution will swell until they eventually burst, known as **hemolysis**.

Figure 8-21 The effect on red blood cells placed into solutions of different tonicity. (a) An isotonic solution: no net flow of water; (b) a hypotonic solution: water flows into the cell; (c) a hypertonic solution: water flows out of the cell. [David M. Phillips/Science Source]

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Math Tips: These margin notes provide students with a math refresher at relevant points in the text.

Figure 8-16 The most common intravenous (IV) solution is physiological normal saline, which has a sodium chloride concentration of 0.90%. [dtimiraos/Getty Images]

MATH TIP: Remember, the symbol % means "per 100," or "divided by 100," or "for every 100." For example, 12% means $\frac{12}{100}$ or 12 out of 100.

% Mass/Volume Concentration

Another concentration unit similar to mass/volume (m/v) is percent mass/volume (% m/v). To calculate a **% m/v concentration**, the mass of the solute must be in grams (g) and the volume of solution must be in milliliters (mL). The mass/volume is then calculated and multiplied by 100 to obtain a percentage, as shown in the equation below.

$$\% \text{ m/v} = \frac{\text{g solute}}{\text{mL solution}} \times 100$$

For example, the most common IV solution, physiological normal saline, has a sodium chloride concentration of 0.90% m/v (Figure 8-16), which means 0.90 g of NaCl per 100 mL of solution. Table 8-4 lists some other common IV solutions and the concentration of their solutes, printed on the label as a % mass/volume. These IV solutions are used routinely to treat dehydration and electrolyte imbalances.

Table 8-4 Solute Concentrations of Some Common IV Solutions

Common Name	Solution Concentration
Normal saline (NS)	0.90% NaCl
Half-normal saline (1/2 NS)	0.45% NaCl
D5W	5% dextrose
D5NS	5% dextrose, 0.9% NaCl
Lactated Ringer's (LR)	0.60% NaCl, 0.31% sodium lactate, 0.03% KCl, 0.02% $CaCl_2$

To prepare a solution from a more concentrated solution use the dilution equation below.

$$C_1 \times V_1 = C_2 \times V_2$$

where C_1 = the concentration of the more concentrated solution (stock solution),
V_1 = the volume of the more concentrated solution (stock solution),
C_2 = the concentration of the dilute solution,
V_2 = the volume of the dilute solution.

To use this equation, *three* of the four variables must be given, and then you can solve for the fourth, unknown variable.

When using the solution dilution equation, C_1 and C_2 can be any unit of concentration, provided the units are the same: m/v, % m/v, *M*. Similarly, the volume can be any volume unit, provided the units are the same. For example, if molarity, *M*, is the concentration unit, the solution dilution equation can be rewritten as shown below. The steps for solving a solution dilution calculation are described in the *Guidelines: Solving a Solution Dilution Problem.*

$$M_1 \times V_1 = M_2 \times V_2$$

MATH TIP: Recall from algebra that when you have one equation, you can only solve for one unknown. If you have two equations, then you can solve for two unknowns.

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Guidelines: These boxes throughout the text provide step-by-step instructions for a common type of chemistry exercise (such as calculations or nomenclature).

Guidelines Solving a Solution Dilution Problem

Example: To prepare 100 mL of a 1% aqueous saline solution from a 10% aqueous saline stock solution, what volume of the stock solution should be added to a 100 mL volumetric flask?

Step 1. Begin by identifying the three given variables and the unknown variable from the information in the question. Read the question carefully to identify the three-of-four variables required in order to use the dilution equation. Look for clues that indicate a reference to the more concentrated or stock solution versus the dilute solution.

$C_1 = 10\%$ *concentration of stock solution*
$C_2 = 1\%$ *concentration of dilute solution*
$V_1 = ?$ *volume of the stock solution needed*
$V_2 = 100\text{ mL}$ *volume of the dilute solution*

Step 2. Rearrange the dilution equation algebraically, solving for the unknown variable. Divide both sides of the dilution equation by the variable you want to remove from one side of the equation.

The dilution equation:

$$C_1 \times V_1 = C_2 \times V_2$$

In this example, we want to solve for V_1, the unknown, so we divide both sides by C_1:

$$\frac{\cancel{C_1} \times V_1}{\cancel{C_1}} = \frac{C_2 \times V_2}{C_1}$$

$$V_1 = \frac{C_2 \times V_2}{C_1}$$

Step 3. Substitute the variables identified in step 1 into the algebraically rearranged equation from step 2 and solve for the unknown variable. Substitute the known variables into the rearranged dilution equation. Cancel like units and solve.

$$V_1 = \frac{1\cancel{\%} \times 100\text{ mL}}{10\cancel{\%}} = 10\text{ mL}$$

Transfer 10 mL of the 10% stock solution to a 100 mL volumetric flask and then add water to the fill mark and mix.

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Additional features to help students with problem solving and math skills:

Worked Exercises: The author walks students through the steps for solving problems, pointing out important concepts and explaining each step.

WORKED EXERCISES **Amount of Solute in a Solvent**

8-5 Student A adds 50 g of sodium nitrate, $NaNO_3$, to 100 mL of water and observes a clear colorless solution. Student B adds 100 g of sodium nitrate to 100 mL of water and observes a cloudy heterogeneous mixture. Which student has made a supersaturated solution? Why do these students observe different results even though they are both mixing the same solute and solvent?

Solution:

Student B has prepared a *supersaturated* solution. Student B has the same but twice as much solute as Student A, exceeding the amount that dissolves in 100 mL of water at that temperature, and therefore, resulting in a supersaturated solution.

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Practice Exercises: Additional Practice Exercises follow each set of Worked Exercises, allowing students to test their understanding of the material.

PRACTICE EXERCISES

5 Explain the meaning of the saying "like dissolves like." What "like" properties does this statement refer to? How is sugar "like" water?

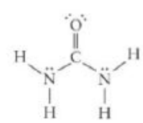
6 Based on the solubility characteristics of solutions, why do kidney stones form?

7 On the macroscopic scale, how does a supersaturated solution look different from a saturated solution? Explain the difference on the molecular level.

8 Explain why each of the following solutes would or would not be expected to dissolve in hexane.

a. NaCl
b. CCl_4
c. CH_3OH
d. pentane (C_5H_{12})

9 Urine, produced by the kidneys, is an aqueous solution containing urea and other small polar molecules and ions. Illustrate how a molecule of urea dissolves in water by showing one molecule of urea hydrogen bonding to three water molecules at three different places in the molecule. (*Hint*: Look for the polar bonds.)



Urea

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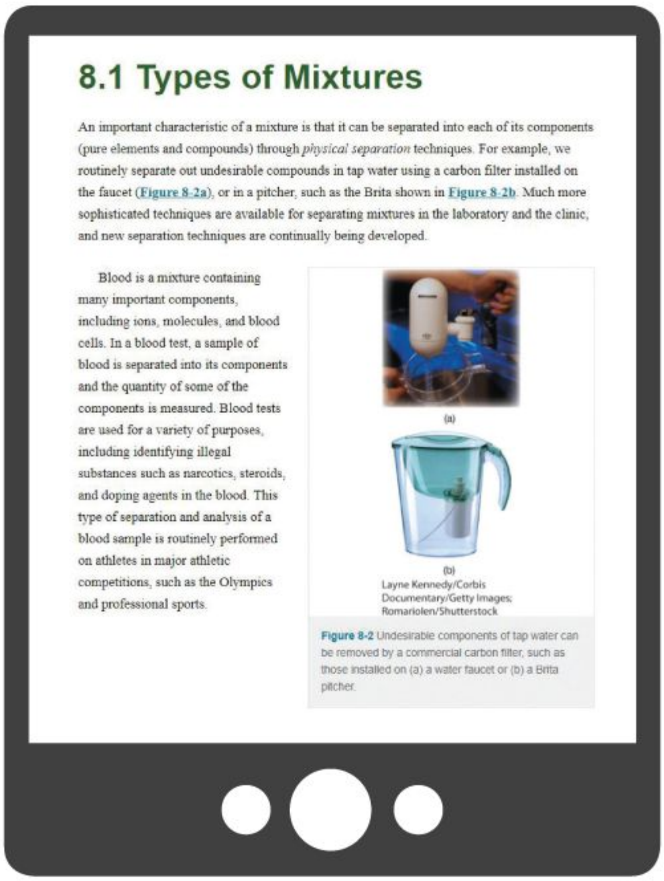
There is also a **Math Appendix** that provides students with a primer of necessary math skills and tips on using the TI-30Xa scientific calculator.

Promotes Independent Practice with SaplingPlus

Guinn's third edition of *Essentials of General, Organic, and Biochemistry* is now paired with SaplingPlus, promoting student study and practice. Adaptive LearningCurve assignments help direct purposeful reading and study, while tutorials and case studies help students synthesize and apply their understanding. In addition, students and instructors will have access to the acclaimed Sapling Learning online homework and access to industry-leading peer-to-peer support for help with implementation and technical support.

SaplingPlus now includes a new VitalSource e-book. This e-book is also available through an app that allows students to read offline or have the book read aloud to them. Additionally, students can highlight, take notes, and search for key words.

Vitalsource



8.1 Types of Mixtures

An important characteristic of a mixture is that it can be separated into each of its components (pure elements and compounds) through *physical separation* techniques. For example, we routinely separate out undesirable compounds in tap water using a carbon filter installed on the faucet (Figure 8-2a), or in a pitcher, such as the Brita shown in Figure 8-2b. Much more sophisticated techniques are available for separating mixtures in the laboratory and the clinic, and new separation techniques are continually being developed.

Blood is a mixture containing many important components, including ions, molecules, and blood cells. In a blood test, a sample of blood is separated into its components and the quantity of some of the components is measured. Blood tests are used for a variety of purposes, including identifying illegal substances such as narcotics, steroids, and doping agents in the blood. This type of separation and analysis of a blood sample is routinely performed on athletes in major athletic competitions, such as the Olympics and professional sports.

(a) (b) Layne Kennedy/Corbis Documentary/Getty Images; Romariolen/Shutterstock

Figure 8-2 Undesirable components of tap water can be removed by a commercial carbon filter, such as those installed on (a) a water faucet or (b) a Brita pitcher.

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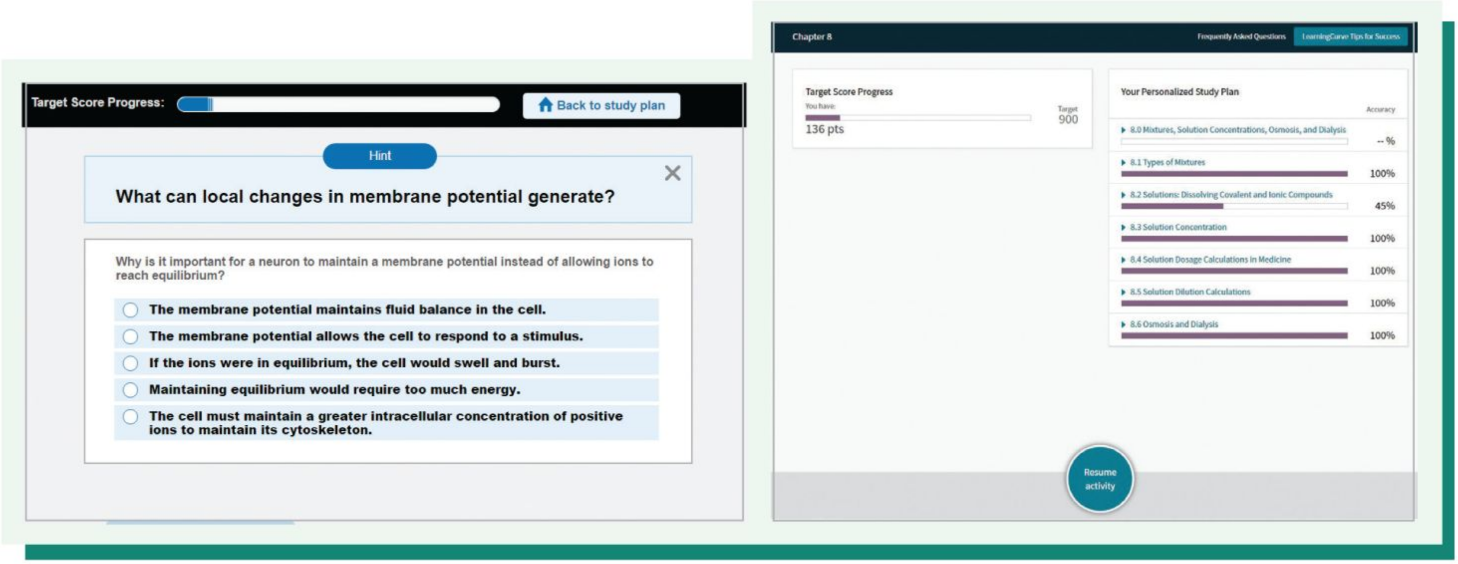
LearningCurve is an adaptive quizzing system that promotes purposeful reading and study. These game-like formative assessments are available for every chapter of the text.



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Tutorial Questions take students through a multi-concept question by breaking it into individual stepped-out problems. These activities help break down complex problems and guide students with hints and targeted feedback at each step, building their confidence and competence. Students can then apply these problem-solving skills when they tackle similar problems on their own.

In addition, **Case Studies** give students the opportunity to build upon their knowledge and synthesize their understanding. These more challenging problems apply chemistry concepts to medical topics.

An order was given to infuse 800. units per hour of the anticoagulant heparin for a patient in the emergency room. The IV bag supplied contains 250. mL of a heparin solution at a concentration of 100. units per milliliter. How many hours will the IV run?

In order to solve this problem, you will want to determine what is being asked.

What quantity do you need to determine to solve this problem?

- ○ flow rate
- ○ units
- ○ molar concentration
- ○ equivalents per liter

An order was given to infuse 800. units per hour of the anticoagulant heparin for a patient in the emergency room. The IV bag supplied contains 250. mL of a heparin solution at a concentration of 100. units per milliliter. How many hours will the IV run?

Since you are determining the flow rate, what is the formula?

flow rate = $\frac{\square}{\square}$

Answer Bank: time | volume | moles | units

An order was given to infuse 800. units per hour of the anticoagulant heparin for a patient in the emergency room. The IV bag supplied contains 250. mL of a heparin solution at a concentration of 100. units per milliliter. How many hours will the IV run?

You determined that the flow rate = $\frac{\text{volume}}{\text{time}}$.

Which conversion factor for the infusion rate will ybe used in the calculation?

- ○ $\frac{1\ \text{h}}{800.\ \text{units}}$
- ○ $\frac{800.\ \text{units}}{1\ \text{h}}$

Which conversion factor for the concentration will be used in the calculation?

- ○ $\frac{100.\ \text{units}}{1\ \text{mL}}$
- ○ $\frac{1\ \text{mL}}{100.\ \text{units}}$

An order was given to infuse 800. units per hour of the anticoagulant heparin for a patient in the emergency room. The IV bag supplied contains 250. mL of a heparin solution at a concentration of 100. units per milliliter. How many hours will the IV run?

You determined that the flow rate = $\frac{\text{volume}}{\text{time}}$. Now organize the conversion factors so that the proper units will be calculated and perform the calculation.

flow rate = $\left(\frac{\square}{\square}\right) \times \left(\frac{\square}{\square}\right) =$ ______ $\frac{\text{mL}}{\text{h}}$

Answer Bank: 1 mL | 1 h | 800. units | 100. units

An order was given to infuse 800. units per hour of the anticoagulant heparin for a patient in the emergency room. The IV bag supplied contains 250. mL of a heparin solution at a concentration of 100. units per milliliter. How many hours will the IV run?

You determined that the flow rate = $\frac{\text{volume}}{\text{time}} = 8.00\ \frac{\text{mL}}{\text{h}}$. Now you can determine the hours the IV will run.

Time = ______ h

Anatomy of a Sapling Problem

Sapling offers multiple question types that enhance student engagement and understanding. Each problem includes hints,answer-specific feedback, and detailed solutions, facilitating student learning and emphasizing the pedagogical value of homework.

Hints attached to every problem encourage critical thinking by providing suggestions for completing the problem, without giving away the answer.

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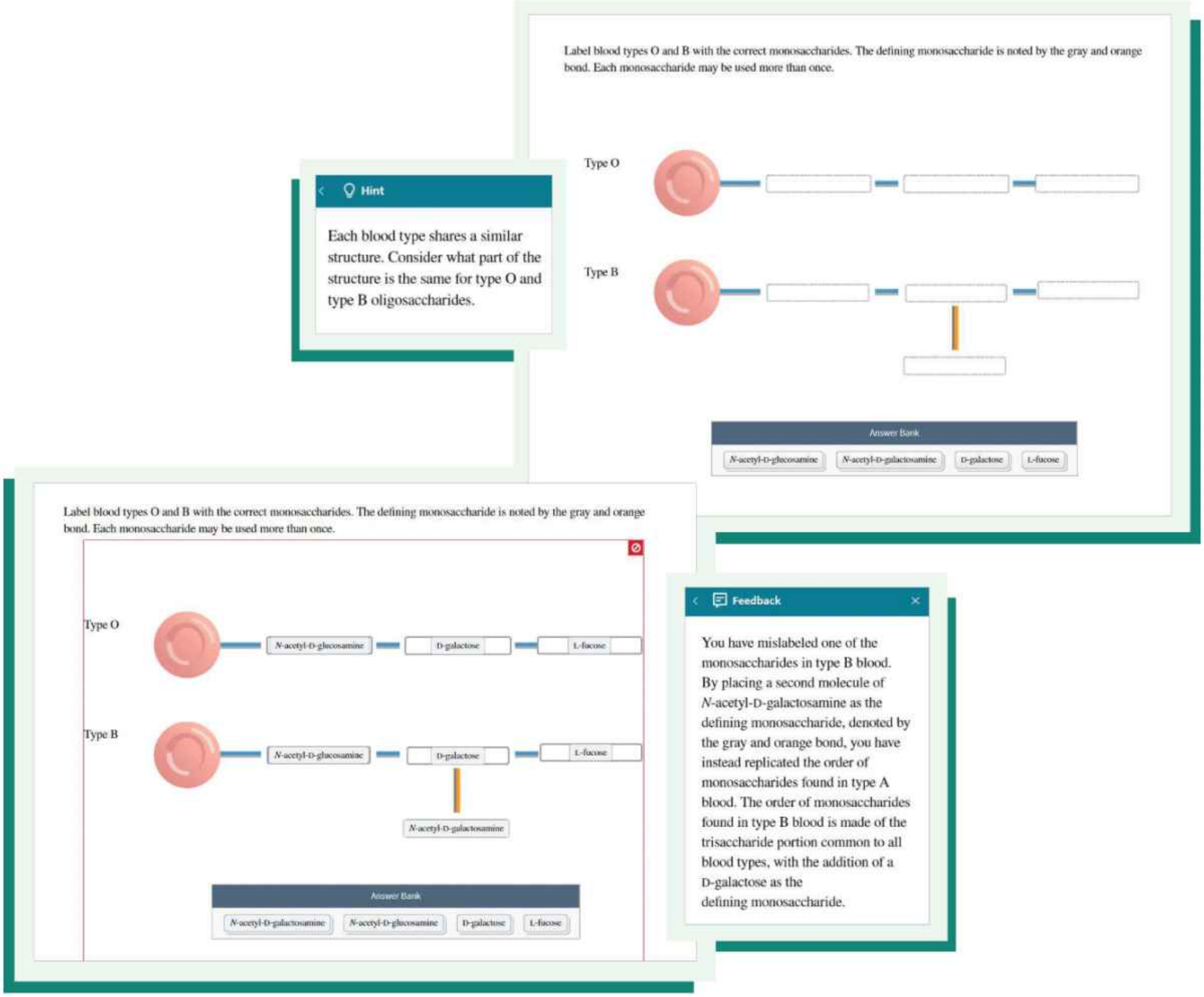

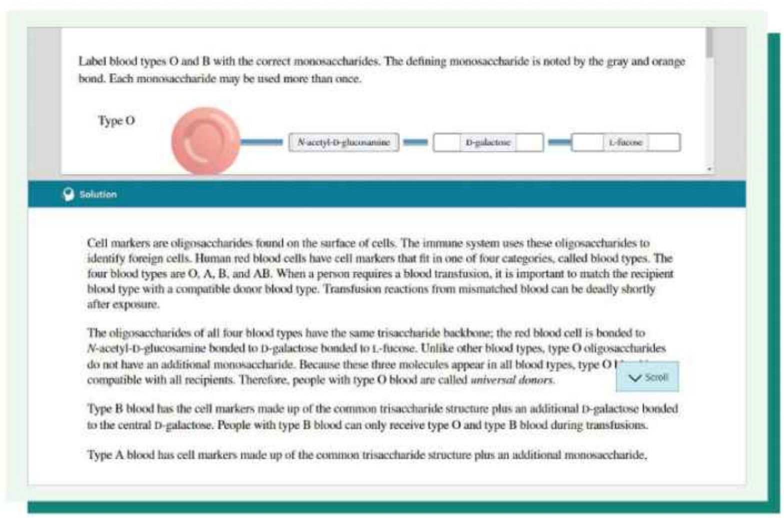

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Selected Updates to the Third Edition

1. Matter, Energy, and Measurement
2. Atomic Structure and Radioisotopes
3. Ionic and Covalent Compounds
4. Molecular Geometry, Polarity, and Intermolecular Forces of Attraction
5. Chemical Quantities and Introduction to Reactions
6. Chemical Reactions: Energy, Rates, and Equilibrium
7. Changes of State and Gas Laws
8. Mixtures, Solution Concentrations, and Diffusion
9. Acids and Bases, pH, and Buffers
10. Introduction to Organic Chemistry: Hydrocarbon Structure
11. Alcohols, Phenols, Thiols, Ethers, and Amines
12. The Carbonyl Containing Functional Groups
13. The Common Organic Reactions in Biochemistry
14. Carbohydrates: Structure and Function
15. Lipids: Structure and Function
16. Proteins: Structure and Function
17. Nucleotides and Nucleic Acids
18. Energy and Metabolism

Another Guidelines Box for using dimensional analysis has been added to chapter 1.

To provide more focus to chapters, the topic of compounds from the previous edition has been divided into two shorter chapters: Chapter 3 "Ionic and Covalent Compounds" and Chapter 4 "Molecular Geometry, Polarity, and Intermolecular Forces of Attraction."

A new section on the basic types of reactions has been added to Chapter 5.

To provide more focus to chapters, the topic of chemical quantities and reactions from the previous edition has been divided into two shorter chapters: Chapter 5 "Chemical Quantities and Introduction to Reactions" and Chapter 6 "Chemical Reactions: Energy, Rates, and Equilibrium."

Chapters 8 and 9 now appear before the introduction of organic chemistry.

The organization of the text has been rearranged so that the four organic chemistry chapters (Chapters 10-13) are together and follow the general chemistry chapters (Chapters 1-9). The chapter on functional groups has been divided into two chapters. (Chapters 11 and 12).

Reorganized coverage of hydrocarbon structure.

Stereoisomers are introduced in this chapter where they are then applied to monosaccharides. A new optional section describing how to convert a Fischer projection to a Haworth projection has been added to Chapter 14.

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Acknowledgments

We are grateful to the many instructors who gave their time and expertise to reviewing revised drafts of chapters in the third edition, some of whom reviewed many chapters of the book.

Jessica Aguirre, *University of Connecticut*

Angela Allen, *Shaw University*

Premilla Arasasingham, *Moorpark College, El Camino College*

Christopher Babayco, *Columbia College*

Yiyan Bai, *Houston Community College*

Lucas Beagle, *Kettering College*

Peter Bell, *University of Pittsburgh*

Stacey-Ann Benjamin, *Broward College*

Terrence Black, *Nassau Community College*

Dina Borysenko, *Milwaukee Area Technical College*

Kristin Butterworth, *East Texas Baptist University*

Gary Campbell, *Northampton Community College*

Bernard Castillo, *University of the Virgin Islands*

Tammy Clark, *Viterbo University*

Stuart Cohen, *Horry-Georgetown Technical College*

Olga Fryszman, *Miramar College*

David Gelormo, *Northampton Community College*

Cynthia Gilley, *San Diego Miramar College*

Pierre Goueth, *MiraCosta College*

Bonnie Hall, *Grand View University*

Abby Hodges, *MidAmerica Nazarene University*

Jason Holland, *University of Central Missouri*

Amber Howerton, *Nevada State College*

Jennifer Jamison, *University of Arkansas-Fort Smith*

Peng Jing, *Purdue University Fort Wayne*

Bret Johnson, *The College of St. Scholastica*

Matthew Johnston, *Lewis-Clark State College*

Kristopher Keuseman, *Mount Mercy University*

Amy Kovach, *Roberts Wesleyan College*

Mary Lamar, *Eastern Kentucky University*

Rebecca LaRue, *Clermont College at University of Cincinnati*

William Loffredo, *East Stroudsburg University*

William Magilton, *Northampton Community College*

Brian Martinelli, *Nevada State College*

Christopher Massone, *Molloy College*

Melissa McAlexander, *Notre Dame de Namur University*

Shabana Meyering, *Northern Virginia Community College*

Margaret Ndinguri, *Eastern Kentucky University*

Loralee Ohrtman, *Lewis-Clark State College*

James Patrone, *Rollins College*

Julie Peller, *Valparaiso University*

Tchao Podona, *Miami Dade College*

Uttam Pokharel, *Nicholls State University*

Elizabeth Pulliam, *Tallahassee Community College*

Kazi Rahman, *University of Mount Olive*

Bhavna Rawal, *Houston Community College*

Tanea Reed, *Eastern Kentucky University*

Billy Samulak, *Fitchburg State University*

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Lisa Sharpe *Elles, University of Kansas*

Wendy Shuttleworth, *Lewis-Clark State College*

Joseph Simard, *University of New England*

Kent Strodtman, *Columbia College*

Harishchandra Subedi, *Western Nebraska Community College*

Ann Taylor, *Wabash College*

Harold Trimm, *SUNY Broome*

Linda Waldman, *Cerritos College*

Sen Wang, *California State University Dominguez Hills*

Xin Wen, *Cal State Univ Los Angeles*

Bill Williams, *Hudson Valley Community College*

Mustafa Yatin, *Salem State University*

We are also grateful to the many reviewers whose feedback helped us broadly shape the third edition of this text.

Brian Abbott, *The University of Arkansas-Fort Smith*

Sherri Adams, *Gillette College, NWCCD*

Jessica Aguirre, *University of Connecticut*

Angela Allen, *Shaw University*

Valerie Barber, *University of Alaska*

Stacey-Ann Benjamin, *Broward College*

Lidia Berbeci, *Broward College*

Kristin Butterworth, *East Texas Baptist University*

Gary Campbell, *Northampton Community College*

Ken Capps, *College of Central Florida*

Gerard Caprio, *DeSales University*

Liheng Chen, *Aquinas College*

Emma Chow, *Palm Beach State College*

Kristin Clark, *Ventura College*

Tammy Clark, *Viterbo University*

Stuart Cohen, *Horry-Georgetown Technical College*

Timothy Dransfield, *UMass Boston*

GC Flowers, *Albany State University West Campus*

Peter Friedman, *Grossmont College*

Tom Gardner, *Saint Cloud State University*

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Cynthia Gilley, *San Diego Miramar College*

Isabelle Haithcox, *Notre Dame de Namur University*

Bonnie Hall, *Grand View University*

Sangita Handa, *School of Liberal Arts & Sciences, Lafayette Ivy Tech*

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Heather Hollandsworth, *Harding University*

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Shelli Hull, *Tarrant County College South Campus*

Michael Hydorn, *Germanna Community College*

Jennifer Jamison, *University of Arkansas-Fort Smith*

Peng Jing, *Purdue University Fort Wayne*

Bret Johnson, *The College of St. Scholastica*

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Bryan Schmidt, *Minot State University*

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Robert Shapiro, *Becker College*

Lisa Sharpe *Elles, University of Kansas*

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Wendy Shuttleworth, *Lewis-Clark State College*

Joseph Simard, *University of New England*

Cynthia Simmons, *North Lake College*

Julianne Smist, *Springfield College*

Duane Smith, *Nicholls State University*

Patricia Snyder, *Florida Atlantic University*

Allison Soult, *University of Kentucky*

Koni Stone, *CSU, Stanislaus*

Monica Strada, *Sacred Heart University*

Amy Taketomo, *Hartnell College*

Ann Taylor, *Wabash College*

Theunis van *Aardt, West Liberty University*

Maria Vogt, *Bloomfield College*

Linda Waldman, *Cerritos College*

This edition comes with an updated media package, including new Sapling homework questions, tutorials, and case studies. We thank our media review participants for their help in re-envisioning the media package.

Jessica Aguirre, *University of Connecticut*

Premilla Arasasingham, *Moorpark College, El Camino College*

Yiyan Bai, *Houston Community College*

PJ Ball, *Northern Kentucky University*

Lucas Beagle, *Kettering College*

Peter Bell, *University of Pittsburgh*

Ken Capps, *College of Central Florida*

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Teresa Cowan, *Baker College of Auburn Hills*

Caitlin Deskins, *Quincy University*

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Christian Schaefer, *University of Massachusetts Boston*

Timothy Searls, *University of Southern Maine*

Joseph Simard, *University of New England*

Chad Snyder, *Grace College*

David Soriano, *University of Pittsburgh-Bradford*

And finally, we thank our faculty advisory board who provided valuable and extensive feedback on many aspects of the third edition, including the organization and media content.

Stacey-Ann Benjamin, *Broward College*

Tammy Clark, *Viterbo University*

Bonnie Hall, *Grand View University*

Michael Hydorn, *Germanna Community College*

Jennifer Jamison, *University of Arkansas-Fort Smith*

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Billy Samulak, *Fitchburg State University*

James Selzler, *Ventura College*

Lisa Sharpe *Elles, University of Kansas*

This textbook would not have been possible were it not for the exceptional dedication provided by the editorial team at W. H. Freeman and Macmillan Learning. My deepest gratitude goes to Maria Lokshin, the development editor on this edition (as well as Randi Rossignol, the senior development editor on the second edition, and Susan Moran, the senior development editor for the first edition). Their commitment was instrumental in achieving the text you see before you today. It has been a great pleasure undertaking a creative endeavor in collaboration with individuals that have such considerable talent and insight. I also want to thank the Senior Program Manager Beth Cole, who deftly managed the many facets of producing a modern textbook: ancillaries, media package, and all the associated activities that must be in place for a textbook to come to fruition. Finally, I appreciate the considerable experience brought to the project by the team that worked on the second edition: acquisitions editors Jessica Fiorillo, Anthony Palmiotto, and Bill Minick.

This textbook benefits from the dedicated efforts of our supplements authors, to whom we extend our most heartfelt gratitude. Rachel Lum prepared the on-line solutions guide that accompanies the more than 1,500 questions that appear in the book. She also revised and wrote the new exercises that appear at the end of each chapter. Mildred Hall created the comprehensive PowerPoint slides for this edition, and revised the iClicker questions. Rachel Jameton and Mildred Hall authored the test bank for this edition. Billy Samulak and Kent

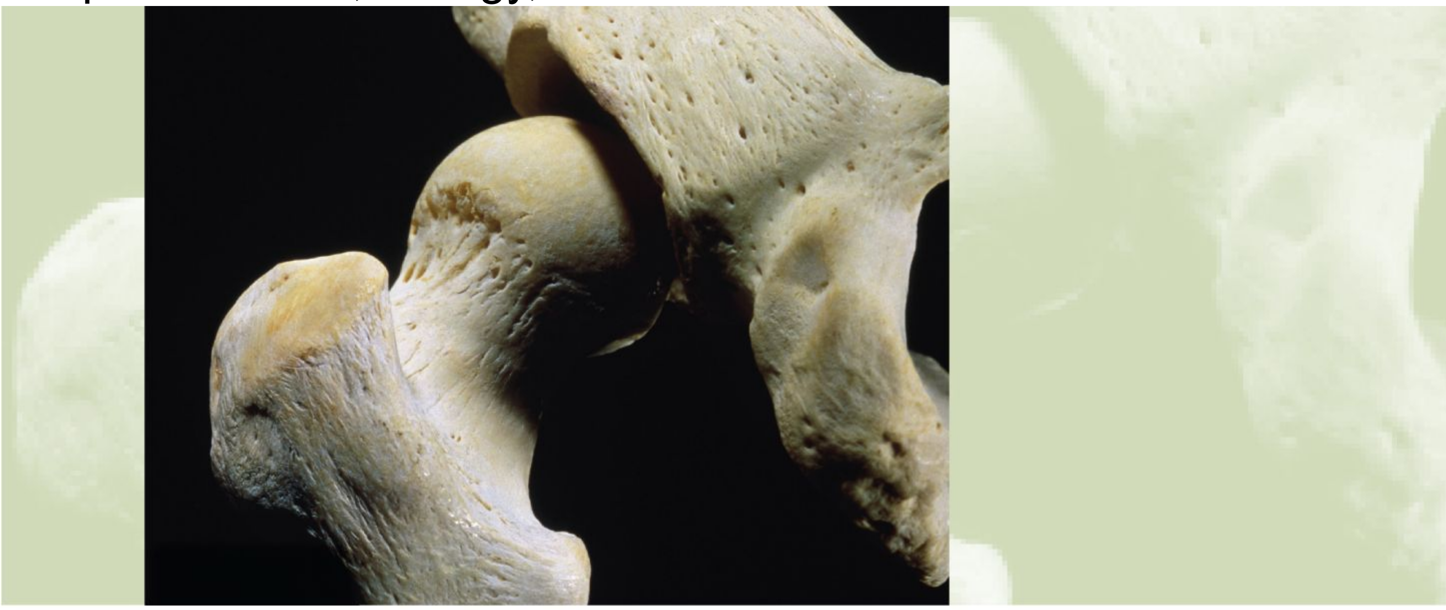
Strodtman wrote the case studies. We would also like to thank Stephanie Ryan and Chad Snyder for authoring and revising the LearningCurve questions for this text.

We extend our appreciation to the staff at Lumina, and to the many essential W. H. Freeman production staff members including senior project content manager Kerry O'Shaughnessy, project manager Linda DeMasi, and associate program manager Aravinda Doss, who carefully shepherded the book through the proof stages; Karen Misler, the author schedule project manager; Beth Rosato, the copyeditor for this edition and the previous edition; Blake Logan and Patrice Sheridan, who created the elegant design; Emiko Paul and Quade Paul (Echo Medical Media) for creating the cover art; Paul Rohloff, who arranged the typesetting and printing; Krystyna Borgen, who researched many of the photographs; and Robin Fadool, photo editor. We owe a special thanks to Kristen Ford, Stacy Benson, Lily Huang, and Kris Hiebner for ably guiding the development of the impressive print and online set of resources available on our book's website. Finally, we thank Maureen Rachford, marketing manager, and the entire sales force for all of their enthusiasm and support.

The pedagogy of the book is greatly enhanced by the artwork produced by Network Graphics. This text is further enhanced and unique in the number and quality of authentic protein structures, which have been rendered by Gregory Williams, who also produced the electron density models. We would also like to thank Alex Panov and Lily Huang for producing the ball-and-stick models and space-filling models included throughout the text.

A special debt of gratitude goes to photographer Catherine Bausinger, who shot the more than 50 new photographs that appeared in the second edition and again in the third edition, Charles Guinn who revised the much-improved math appendix for the second edition, Scott Guinn who served as my editorial assistant throughout the entire third edition revision, and my brother Peter Vogel for all his technology advice. I would also like to express my deepest gratitude to Rebecca Brewer for collaborating with me on the writing and development of the first edition. I also want to thank my colleagues who throughout the writing of the third edition supported my efforts: Rachel Jameton (Lewis-Clarke State College), Rachel Lum, Madeline Mignone (Dominican College), Kimberley Waldron (Regis University), Steve Cartier (Amherst College); and my colleagues at The College of New Rochelle—Richard Thompson, Rebecca Lafleur, Lee Warren, Terry Colarusso, Melanie Harasym, Lynn Petrullo, Faith Kostel-Hughes, and Michael Gilliam.

Chapter 1 Matter, Energy, and Measurement



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Contents

CONCEPTS IN CONTEXT: Osteoporosis and Measurement of Bone Density

It is estimated that one in five American women over the age of 50 has osteoporosis and that more than half of these women will break a bone at some point in their lives. Osteoporosis is a condition characterized by the progressive loss of bone density and creates a greater risk of bone fractures. Most bone fractures in people with osteoporosis occur in the hip, wrist, or spine. A stumble or even a cough is sometimes enough to cause a fracture.

As with many chronic medical conditions, early intervention can limit the progression of the condition. Individuals at risk of osteoporosis are advised to get periodic bone mineral density measurements. The most common way to assess bone strength is with a *b*one *m*ineral *d*ensity (BMD) measurement. The average BMD for an adult is $1.5\ \text{g/cm}^3$, read as "one point five grams per centimeter cubed." Compare this value to the density of some familiar materials:

oak:	$0.72\ \text{g/cm}^3$
carbon fiber:	$1.6\ \ \text{g/cm}^3$
diamond:	$3.3\ \ \text{g/cm}^3$

The greater the BMD, the stronger the bone. Since stronger bone is better able to withstand stress, it is less likely to fracture.

A DEXA (*d*ual *e*nergy *x*-ray *a*bsorptiometry) scan is used to estimate BMD in patients and screen for osteoporosis. *Density* can be calculated from the measured mass of a sample of material divided by its measured volume: $d = \dfrac{m}{V}$. A DEXA scan uses low dose x-rays to measure the *mass* and estimate the *volume* of a section of bone, usually in the hip, wrist, or lumbar spine, from which the BMD is then calculated. A patient's BMD is then compared to the average BMD for healthy young adults of the same gender and ethnicity, and a *T-score* is assigned based on how much the patient's BMD deviates from this average (Figure 1-1).

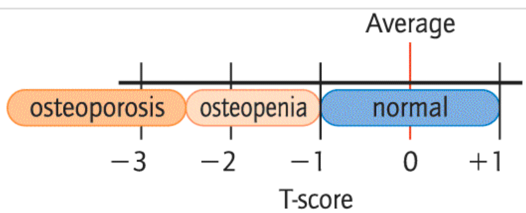

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Figure 1-1 A T-score is assigned to a patient by comparing their BMD (bone mineral density) measurement to the average value for healthy young adults of the same gender and ethnicity.

From a patient's T-score, the doctor has the information needed to make a recommendation for a course of action. Although there is no cure for osteoporosis, lifestyle changes can be made, and medications may be prescribed that can slow or stop the progression of bone mineral loss. Lifestyle changes generally include adding foods to the diet that are high in calcium and vitamin D—helps the body absorb calcium—and doing regular weight-bearing exercise, which stimulates bone growth. In patients with osteoporosis or a history of fractures, one of several medications used to treat osteoporosis may be prescribed. A timely measurement of bone mineral density can inform the best course of treatment and help patients avoid or minimize the more debilitating outcomes of osteoporosis. ●

In this chapter, we focus on measurement and the important role it plays in medicine and science. You will learn how to properly report a measurement, the common units of measurement, and how to perform calculations involving measurements. If you are working in the health care field, you will take measurements routinely and perform calculations that are critical to the health of your patients, such as dosage calculations. Developing the skills learned in this chapter is a basic part of your training and also central to understanding chemistry. But first, we introduce matter and the important concept of energy, which are at the foundation of chemistry.

1.1 Matter and Energy

Chemistry is the study of matter, changes in matter, and the energy associated with those changes. In science, **matter** is defined as anything that has mass and takes up space. Therefore, matter is all the "stuff" around you and in you. We often refer to a specific type of matter—such as aspirin, blood, or air—as a **substance**.

Matter is found in three **states**: solid (s), liquid (l), and gas (g). Examples of all three states of matter are found in the body. The air we breathe is in the gas state, making it possible to quickly fill the lungs with each breath. Blood is in the liquid state, so it is readily pumped by the heart throughout the circulatory system. Skin and bone are in the solid state, providing structural integrity to the body.

In chemistry, we study matter and changes in matter at the **atomic scale**—that which we cannot see—so that we are better able to understand what we observe on the **macroscopic scale**—that which we can see with the naked eye. This includes the matter that makes up the human body, as most diseases can be explained by some malfunction at the atomic level.

We can explain the macroscopic differences we observe in the three states of matter, illustrated in the top part of **Figure 1-2**, by understanding the differences in the states of matter on the atomic scale, illustrated in the bottom part of Figure 1-2.

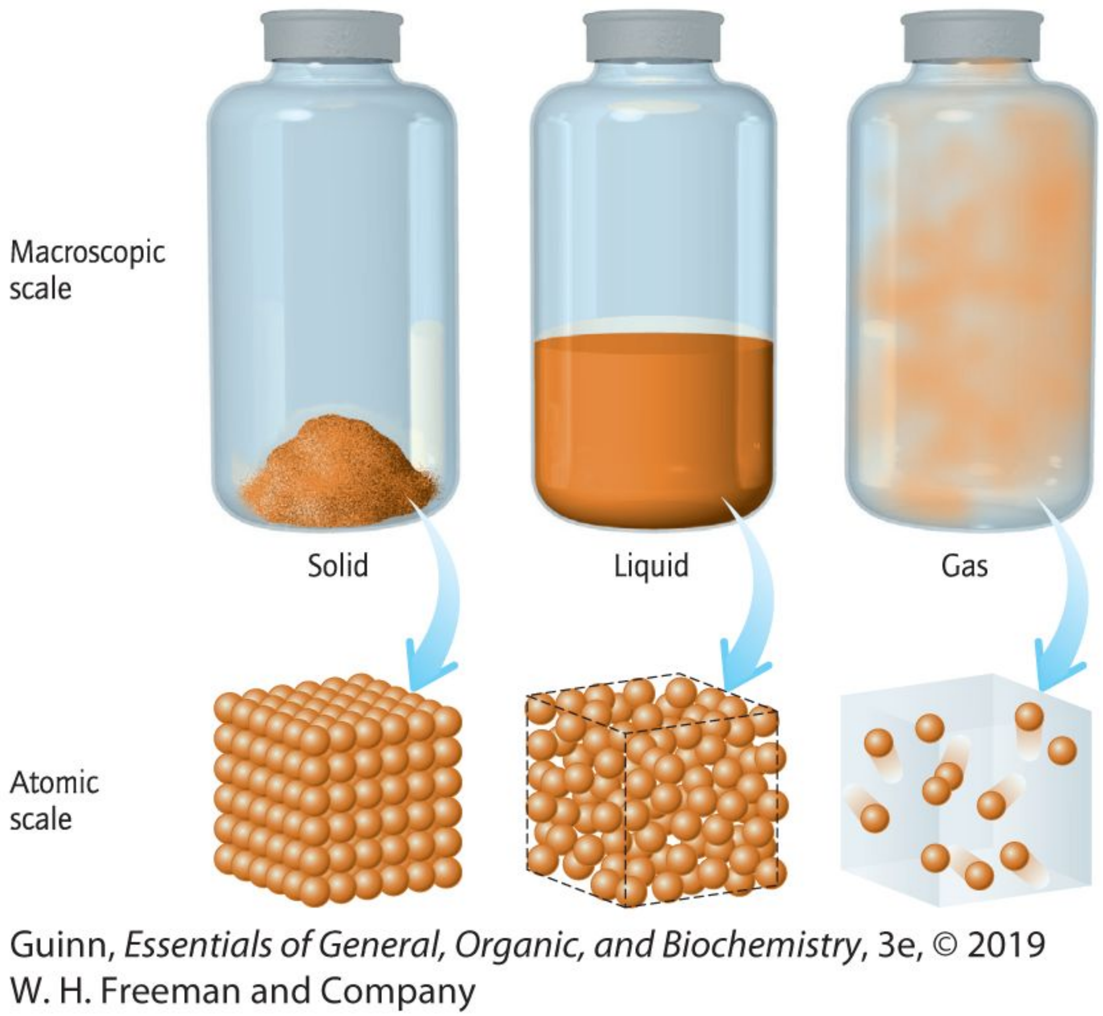


Figure 1-2 Differences observed in the macroscopic properties of solids, liquids, and gases are a result of differences in the way the particles of matter interact on the atomic scale.

Experimental evidence shows that matter is composed of particles. In chapters 2 and 3 you will learn about the nature of these particles. In the **solid** state, particles of matter are close together and in a very ordered arrangement. In the **liquid** state, particles of matter are disordered and farther apart but still interacting with other particles. In the **gas** state,

CORE CONCEPT:

- In the solid state, particles of matter are in an ordered arrangement and interacting.
- In the liquid state, particles are disordered but interacting with other particles.
- In the gas state, particles are far apart and not interacting.

particles of matter are very far apart, interacting only during the occasional collision.

Kinetic and Potential Energy

Hongqi Zang/Alamy

Figure 1-3 The nurse pushing the wheelchair is doing work: the act of moving an object against an opposing force.

Particles of matter are not stationary. The particles of a solid are vibrating; in the liquid state they are moving randomly and tumbling over one another. This is why liquids flow when poured. The particles in a gas are moving randomly and rapidly, and occasionally colliding with the walls of their container and other particles.

The speed at which particles of matter are moving depends on their energy. Hence, matter and energy have an important relationship. In science, **energy** is defined as the capacity to do work, and **work** is defined as the act of moving an object against an opposing force. According to this definition, the nurse in **Figure 1-3** is doing "work" because he is moving an object, the patient in the wheelchair, against an opposing force (gravity and friction).

Kinetic Energy

Energy has two basic forms: *kinetic* energy and *potential* energy. **Kinetic energy** is the energy of *motion*. Kinetic energy applies to both a moving object—the macroscopic scale, and to moving particles—the atomic scale. For example, a car moving at $90\ \text{mph}$ has more kinetic energy than a car moving at $20\ \text{mph.}$ Similarly, a substance has kinetic energy as a result of the motion of its particles. The greater the average speed of its particles, the greater the kinetic energy of the substance.

CORE CONCEPT: A faster moving object has more kinetic energy than a slower moving object.

CORE CONCEPT: For a given substance, particles in the gas state have more kinetic energy than particles in the liquid state, which have more kinetic energy than particles in the solid state.

For a given substance, its particles are moving faster when in the gas state than in the liquid state, and moving faster in the liquid state than in the solid state. For example, steam (water particles in the gas state) has more kinetic energy than liquid water (water particles in the liquid state) because the particles in the gas state are moving faster than the particles in the liquid state.

Temperature is a measure of the average kinetic energy of particles of matter. A substance that has particles with a greater average kinetic energy will have a higher temperature. As particles of matter move faster, their kinetic energy increases, and we note a temperature

CORE CONCEPT: Temperature is a measure of the average kinetic energy of particles of matter. For a given substance, the gas state has a higher temperature than the liquid state, and the liquid state has a higher temperature than the solid state.

rise. Rub your hands together quickly and vigorously. Do you feel them getting warmer? Your hands feel warmer because you have made the particles on the top layer of your skin move faster, increasing their kinetic energy.

CORE CONCEPT: Heat is the kinetic energy transferred from matter at a higher temperature to matter at a lower temperature.

Heat is kinetic energy *transferred* from a sample of matter at a higher temperature to a sample of matter at a lower temperature. Heat always flows from hot to cold. Note that temperature is *not* the same as heat: Temperature is a measure of the average kinetic energy of a sample of matter, whereas heat is the kinetic energy transferred.

Consider what happens on the atomic level when an ice cube is added to a hot cup of coffee. Upon contact with the ice cube, the faster moving particles in the hot coffee transfer some of their kinetic energy to the slower moving particles in the ice (water particles in the solid state). The average kinetic energy of the ice particles increases as a result, and the average kinetic energy of the coffee particles decreases, until both the ice and the coffee particles have the same kinetic energy. On the macroscopic scale, we observe the ice melting (a change from the solid to the liquid state), due to an increase in the kinetic energy of the ice; the coffee becomes cooler, due to a decrease in kinetic energy of the coffee, as measured by its lower temperature.

Alex Livesey/Getty Images

Figure 1-4 The application of an ice pack to the site of an injury causes blood vessels to constrict by reducing their kinetic energy.

Athletic trainers, or anyone administering first aid for a sprained ankle or pulled muscle, will usually apply an ice pack to the injury (Figure 1-4). Sprains, tears, and pulled muscles typically cause swelling and pain because fluid leaks from nearby blood vessels. Applying an ice pack to the site of the injury causes the blood vessels to constrict so less fluid leaks from the blood vessels. Blood vessels constrict because the ice pack has less kinetic energy (lower temperature) than the body, so heat (kinetic energy) is transferred from the site of the injury to the ice, causing a decrease in the kinetic energy of the nearby blood vessels. A hot pack is then often applied a few days later. Since the hot pack has more kinetic energy (higher temperature) than the body, heat (kinetic energy) is transferred from the hot pack to the joint or muscle, causing blood vessels in the area to dilate, which increases blood flow and helps restore movement to the joint.

Potential Energy

CORE CONCEPT: Matter has potential energy as a result of position or composition. Some substances are higher in potential energy (less stable) while other substances are lower in potential energy (more stable)

Potential energy is *stored* energy. Matter has potential energy as a result of its *position* or *composition*. A ball poised at the top of a precipice, for example, has high potential energy as a result of its position—we say it is *unstable* (**Figure 1-5**). As the ball falls, its potential energy is converted into kinetic energy—the energy of motion. After it comes to rest at the bottom of the hill, it has low potential energy—we say it is *stable.*

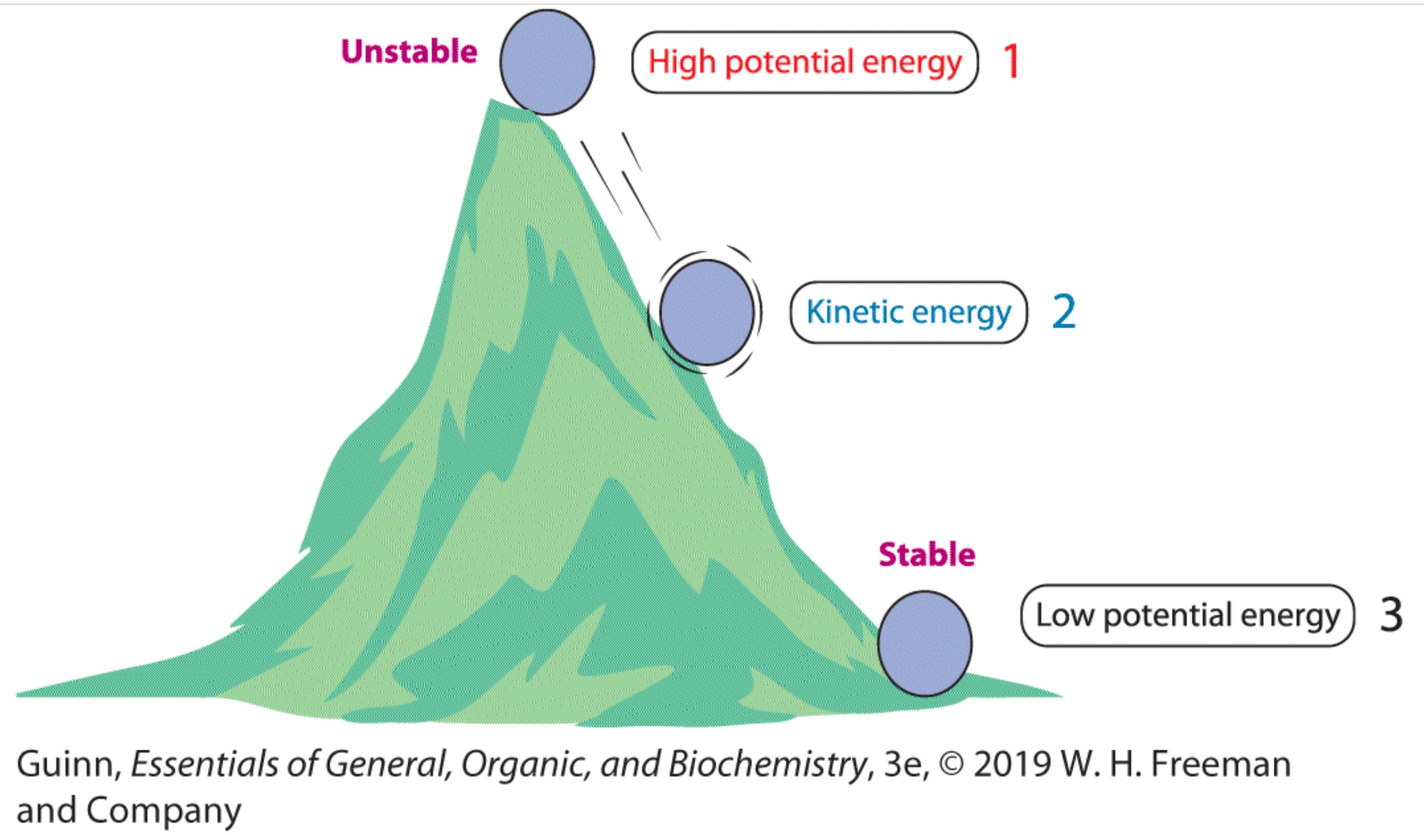


Figure 1-5 (1) The ball at the top of the hill has higher potential energy (it is *unstable*); (2) as the ball falls its potential energy is converted into kinetic energy; (3) at the bottom of the hill, the ball has lower potential energy (it is *stable*).

Matter has potential energy as a result of its *composition*—the nature of the particles of matter. Carbohydrates, fats, and fuels are valued for their *high* potential energy. For example, the propane used in a barbeque grill has high potential energy that is converted into heat (kinetic energy) upon ignition, in a chemical reaction with the oxygen in the air (**Figure 1-6**). You use the heat produced to cook your food. Heat is released because the reaction of propane and oxygen produces substances that are lower in potential energy—carbon dioxide and water—substances more stable than propane and oxygen. We will learn about chemical reactions like this one in chapter 6.

WORKED EXERCISES

Matter and Energy

1-1 For a given substance, in which state of matter will the particles be the farthest apart from one another and why?

a. solid
b. liquid

c. g
a

tab62/Shutterstock.com

Figure 1-6 The heat from a barbeque grill results from the reaction of propane and oxygen, which are converted into the lower potential energy products carbon dioxide and water.

s

Solution:

Answer c. Particles in the gas state are the farthest apart because they have the greatest kinetic energy.

1-2 Indicate the type of energy represented in each of the following examples: *kinetic* energy or *potential* energy.

a. a man diving from a diving board
b. a woman standing on the edge of a diving board
c. the oatmeal you had for breakfast

Solution:

a. Kinetic energy because the diver is in motion

b. Potential energy because the diver is in an unstable position on the edge of the diving board. When she jumps, her potential energy will be converted to kinetic energy and when she lands, she will be in a more stable position.

c. Potential energy because of the chemical composition of the oatmeal. Oatmeal is a carbohydrate (high potential energy) that can be converted to energy and compounds with lower potential energy.

1-3 If you add a hot block of aluminum to a beaker of water that is at room temperature, will the temperature of the water in the beaker increase or decrease? Explain using the terms *heat*, *temperature*, and *kinetic energy* in your explanation.

Solution:

The temperature of the water in the beaker will increase because *heat* is transferred from the particles in the aluminum block, which have a higher *kinetic energy*, to the water, which has particles with a lower *kinetic energy*. An increase in kinetic energy is measured as a rise in *temperature*. Heat transfer will stop when the aluminum and the water have the same kinetic energy (they are at the same temperature).

PRACTICE EXERCISES

(You can find the answers to the Practice Exercises at the end of the chapter.)

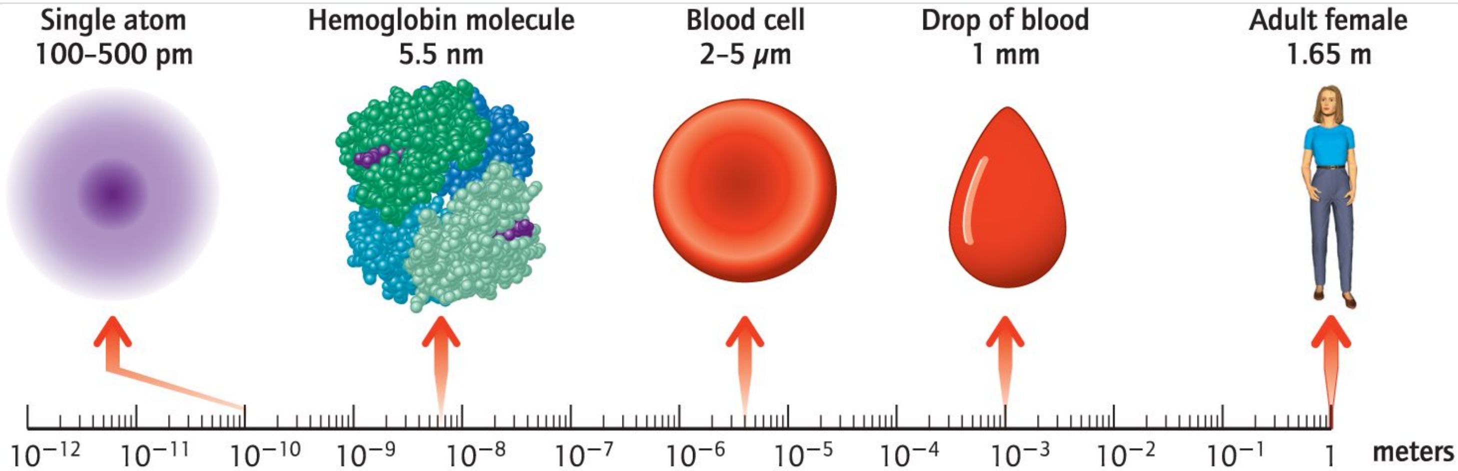
1. Describe two macroscopic differences between water in the liquid state and water in the gas state. Account for these differences by describing how these two states of matter differ on the atomic scale. How is the kinetic energy of the particles for these two states of matter different for this substance? Which has a higher temperature, the liquid phase or the gas phase?

2. Indicate whether *potential energy* or *kinetic energy* is represented in each of the following examples:
   a. a compressed spring
   b. a windmill turning
   c. particles in the gas state colliding with the walls of their container
   d. a skier skiing down a mountain
   e. pasta prepared for a meal

3. Is heat a type of *kinetic energy* or *potential energy*?

1.2 Units and Measurement

Consider a drop of blood on the head of a pin, illustrated in **Figure 1-7**. It is about 1 mm in diameter and you can *see* it with the naked eye—the macroscopic scale. If you look at this droplet of blood through a microscope, you will see that it is composed of millions of red blood cells, each with a diameter 1,000 times smaller than the droplet of blood on the head of a pin—the **microscopic scale**. Imagine—because you cannot see it with a microscope—looking inside one of these red blood cells. You would “see” millions of hemoglobin molecules, the substance that transports oxygen from the lungs to the tissues throughout the body. A hemoglobin molecule has a diameter 1,000 times smaller than a red blood cell and a million times smaller than the droplet of blood. If you imagine zooming in on a molecule of hemoglobin, you will “see” that it is composed of approximately 10,000 atoms, including four iron atoms that are 100 times smaller than a hemoglobin molecule. Molecules and atoms are so small they cannot be observed even with a light microscope—their size is on the atomic scale.



Guinn, *Essentials of General, Organic, and Biochemistry*, 3e, © 2019 W. H. Freeman and Company

Figure 1-7 (left to right) Viewing matter on different scales: The atomic scale (a single atom and a molecule of hemoglobin); the microscopic scale (a red blood cell); and the macroscopic scale (a drop of blood and a woman).

Systems of Measurement

To study matter, we need to be able to measure it. A measurement always includes a numerical value (a number) and a unit. When scientists and medical professionals take a measurement, such as a patient’s weight, height, or temperature, they report both a *number* and a *unit*. The **unit** indicates the type of quantity measured and the system used to measure it. For example, they might report the *weight* of a patient as 112 lb, where lb is the abbreviation for *pounds*, a *unit* of mass in the English system. It would be incorrect to report just the number, 112.

CORE CONCEPT: A measurement consists of a numerical value and a unit.

There are two systems of measurement: the *metric system* and the *English system*, each with its own set of units. The **metric system** is the most widely used system of measurement in the world, while the **English system** is used primarily in the United States. The preferred system of measurement in science and medicine is the metric system. Therefore, a practicing health professional in the United States needs to be familiar with both systems.

For example, the weight of the $112\ \mathrm{lb}$ patient in the example above would have been reported as $50.8\ \mathrm{kg}$ in the metric system, where *kg* is the abbreviation for the *kilogram*, a metric unit of mass. The two measurements are equivalent: $112\ \mathrm{lb} = 50.8\ \mathrm{kg},$ but they have been measured in units from different *systems of measurement*. In section 1.4 you will learn how to convert between English and metric units of mass.

The **i**nternational **s**ystem of units, the **SI system**, was created by an international group of scientists to establish a uniform set of units, selecting one standard metric unit for each quantity of measurement. For example, the SI unit of *mass* is the kilogram (kg). SI units are the official units of science and commerce.

Base Units in the Metric System

> **CORE CONCEPT:** The metric system is built on a set of base units, each denoted by a one-letter symbol.

The metric system is built on a set of **base units**, which represent measurable quantities. The common base units, their one-letter abbreviations shown in parentheses, and the quantity that they measure are given below:

- the meter (m), for length and distance
- the gram (g), for mass and weight
- the liter (L), for volume
- the second (s), for time

Prefixes in the Metric System

> **CORE CONCEPT:** A prefix attached to a base unit increases or decreases the numerical value by a factor of 10^{x} or 10^{-x}.

The metric system employs **prefixes**, each with a unique one-letter symbol, that when preceding a base unit acts as a multiplier to increase the numerical value by a factor of 10^{x} or decrease the numerical value by a factor of 10^{-x}. **Table 1-1** shows the names of the common **metric prefixes**, their abbreviations, and the multiplier they represent (the value of 10^{x} or 10^{-x}). Case (lower case or upper case) matters for prefix abbreviations.

Table 1-1 The Metric Prefixes and the Multipliers They Represent

Prefix	Prefix Symbol	Multiplier in Conventional Notation	Multiplier in Scientific Notation	Example Conversion to a Base Unit
Prefixes for quantities larger than the base unit				
tera	T	1 000 000 000 000	10^{12}	$1\ \mathrm{TB} = 10^{12}\ \mathrm{B}$.
giga	G	1 000 000 000		

			10^9	$1\ \text{GB} = 10^9\ \text{B}$*
mega	M	1 000 000	10^6	$1\ \text{MHz} = 10^6\ \text{Hz}$*
kilo	k	1 000	10^3	$1\ \text{km} = 10^3\ \text{m}$
Prefixes for quantities smaller than the base unit				
deci	d	0.1	10^{-1}	$1\ \text{dL} = 10^{-1}\ \text{L}$
centi	c	0.01	10^{-2}	$1\ \text{cm} = 10^{-2}\ \text{m}$
milli	m	0.001	10^{-3}	$1\ \text{mg} = 10^{-3}\ \text{g}$
micro	μ	0.000 001	10^{-6}	$1\ \mu\text{s} = 10^{-6}\ \text{s}$*
nano	n	0.000 000 001	10^{-9}	$1\ \text{nm} = 10^{-9}\ \text{m}$
pico	p	0.000 000 000 001	10^{-12}	$1\ \text{pm} = 10^{-12}\ \text{m}$

*Base units: Hz = hertz; B = byte; s = seconds

MATH TIP:

- A number with a negative exponent is less than 1; a number with a positive exponent is greater than 1. For example, $10^{-2} = 0.01$ and $10^2 = 100.$
- 10^x means that 10 is multiplied by itself *x* times. So 10^3 means 10 is multiplied by itself 3 times: $10 \times 10 \times 10 = 1{,}000.$ Not: $10 \times 3 = 30.$
- 10^{-x} means $\frac{1}{10}$ or 0.1 is multiplied by itself *x* times. So 10^{-3} means $0.1 \times 0.1 \times 0.1 = 0.001$
- Remember: $\frac{1}{10^3} = 10^{-3}.$

Note that the prefixes at the top of the table (kilo, mega, giga, and tera), *increase* the size of the numerical value because the multipliers are *greater* than 1. For example, *kilo*, abbreviated "k," stands for the multiplier $1{,}000\left(10^3\right).$ When this prefix precedes the base unit of mass, the gram (g), it means $1\ \text{kg} = 1{,}000\ \text{g}$, or written in scientific notation,

$1\ \mathrm{kg} = 10^3\ \mathrm{g}$. This type of mathematical equality between two units is known as a **conversion**. In this case, the conversion equates $1\ \mathrm{kg}$ to $1{,}000\ \mathrm{g}$.

MATH TIP: The number 1,000 is written in scientific notation as 1×10^3, where "1" is known as the coefficient and "3" is known as the exponent. When the coefficient is 1, it can be written as 10^3.

CORE CONCEPT: To create a conversion between a prefixed unit and its base unit, set the prefixed unit equal to the multiplier times the base unit: prefixed $\text{unit} = (10^{x \text{ or } -x}) \times (\text{base unit})$.

The prefixes at the bottom of the table (deci, centi, milli, micro, nano, and pico) *decrease* the size of the numerical value because the multipliers are *less* than 1. For example, the prefix *micro*, abbreviated with the Greek letter μ (pronounced "myou"), represents the multiplier $0.000\ 001\ (10^{-6})$; therefore, the conversion between the prefixed unit and the base unit is $1\ \mu\mathrm{g} = 10^{-6}\ \mathrm{g}$.

A review of scientific notation can be found on pages A-6 and A-7 of *Appendix A: Basic Math Review with Guidance for Using a TI-30Xa Scientific Calculator.*

MATH TIP:

Use the EE or EXP key on your calculator to input a number in scientific notation into the calculator:

- type the value of the coefficient, even if it is 1
- press the EE or EXP button
- type the value and sign, if negative, of the exponent using the +/– key
  (Do NOT press ×)

Metric units with prefixes are used for quantities larger or smaller than the base unit to avoid numerical values with many zeros. To see the effect of prefixes, compare the numerical values reported in the metric base unit to an appropriate metric unit with a prefix in the list below:

Item Measured	Using Metric Base Unit	Using Metric Unit with Prefix
Diameter of a drop of blood on the head of a pin	$0.001\ \mathrm{m}$	$1\ \mathrm{mm}$
Diameter of a blood cell	$0.000\ 001\ \mathrm{m}$	$1\ \mu\mathrm{m}$
Memory on a hard drive	$1{,}000{,}000{,}000\ \mathrm{B}$	$1\ \mathrm{TB}$

The prefixes in the metric system can be used with any of the base units. For example, the prefix *milli*, abbreviated "m," represents the multiplier $0.001\ (10^{-3})$; therefore, the prefix "m" preceding any base unit, indicates multiplication by 0.001. For example:

$$1\ \text{mg} = 0.001\ \text{g};$$
$$1\ \text{mm} = 0.001\ \text{m};$$
$$1\ \text{mL} = 0.001\ \text{L}.$$

Texas Instruments

WORKED EXERCISE

Writing Conversions between Metric Base Units and Prefixed Metric Units

1-4 Write the conversion between the meter and the nanometer using the symbols for these units.

Solution:

The meter (m) is the *base unit* for length or distance in the metric system and the nanometer (nm) is a *prefixed metric unit* of length. From Table 1-1, we see that the symbol "n" is the abbreviation for *nano*, and represents the multiplier 10^{-9}. To create the conversion, let 10^{-9} stand in for the n in nm:

CORE CONCEPT: A metric prefix can be used with any base unit, and always represents the same multiplier.

$$1\ \text{nm} = 10^{-9}\ \text{m}.$$

PRACTICE EXERCISES

4. Write the conversion between the milliliter and the liter using the appropriate abbreviations. Write the conversion between the deciliter and the liter.

5. Write the conversion between the microgram and the gram. Write the conversion between the micrometer and the meter. Write the conversion between the microsecond and the second.

6. What is the conversion between the meter and each of the following?
   a. millimeter
   b. decimeter
   c. kilometer

Units of Length, Mass, and Volume

Scientists and health care professionals take measurements on a daily basis. In this section, we describe some of the most common units of measurement in both the English and metric systems. Then, in section 1.4 we describe how to convert between these two systems of measurement.

Units of Length

Length is the distance between two points (one dimension). The **meter** is the base unit of length in the metric system and the SI unit of length. A meter is divided into 10 decimeters (dm), and a decimeter is divided into 10 centimeters (cm), and a centimeter is divided into 10 millimeters (mm). The actual size and relationship between a decimeter, a centimeter, and a millimeter is shown in Figure 1-8.

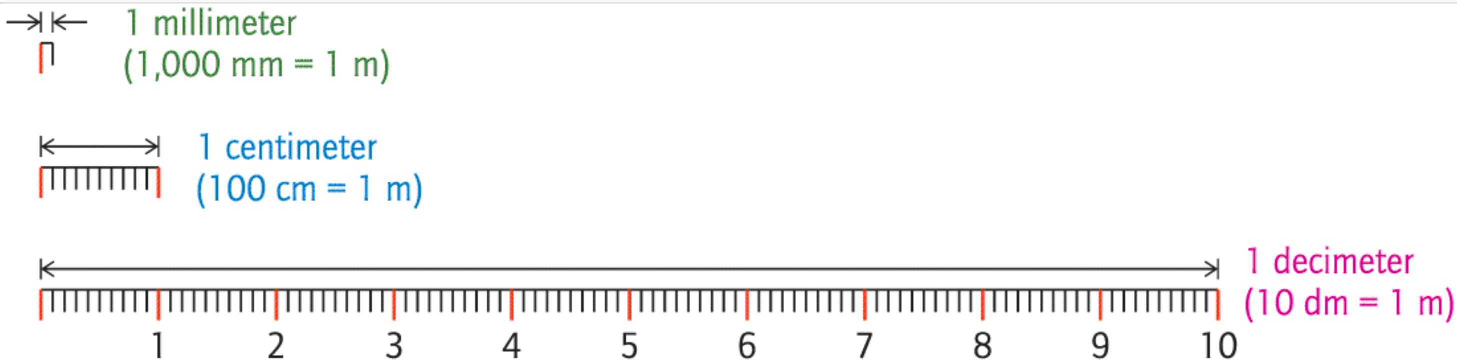


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Figure 1-8 From top to bottom: The actual size and relationship between 1 millimeter $(10^{-3}\ \mathrm{m})$, 1 centimeter $(10^{-2}\ \mathrm{m})$, and 1 decimeter $(10^{-1}\mathrm{m})$ (Not to scale.)

Richard Burkhart/Savannah Morning News/AP Photo

Figure 1-9 Runners preparing to run a 10 km or 5 km road race.

The two examples of length on the macroscale shown in Figure 1-7 are the height of a woman at $1.65\ \mathrm{m}$ and the diameter of a drop of blood on the head of a pin at $1\ \mathrm{mm}.$ A red blood cell has a diameter of $2\text{–}5\ \mu\mathrm{m},$ the example of matter on the microscopic scale. A hemoglobin molecule has a diameter of $5.5\ \mathrm{nm}$ and an iron atom has a diameter of $500\ \mathrm{pm},$ examples of matter on the atomic scale. The kilometer (km) is the only prefixed metric unit used to report distances greater than a meter as for example a $10\ \mathrm{km}$ or $5\ \mathrm{km}$ road race (**Figure 1-9**)

MATH TIP: Note that the conversion $1\ \text{mm} = 10^{-3}\ \text{m}$ (from Table 1-1) is mathematically equivalent to the conversion $10^3\ \text{mm} = 1\text{m}$ both sides by 10^{-3}.

The common units of length in the English system include the inch (in.), the foot (ft), the yard (yd), and the mile (mi). The conversions between some common English units of length are listed below.

$$\begin{aligned} 1\ \text{foot (ft)} &= 12\ \text{inches (in.)} \\ 1\ \text{yard (yd)} &= 3\ \text{feet (ft)} \\ 1\ \text{mile (mi)} &= 5{,}280\ \text{feet (ft)} \end{aligned}$$

In medicine, ultrasound is used as a noninvasive technique to obtain various measurements of length to assess the health and progress of a developing fetus. An ultrasound scan of a fetus, such as the one shown in **Figure 1-10**, can be used to measure the crown-rump length (from the top of the head to the bottom of the buttocks), the biparietal diameter (distance between the sides of the head), the femur (thigh bone) length, and the abdominal circumference. From these measurements, gestational age and growth of the fetus can be estimated. For example, the biparietal diameter of a healthy fetus increases from approximately $2.4\ \text{cm}$ at 13 weeks to $9.5\ \text{cm}$ at term. Structural abnormalities in the fetus, such as spina bifida, can also be diagnosed from these measurements.

Craig Holmes Premium/Alamy

Figure 1-10 An ultrasound scan of a human fetus at 20 weeks. This fetus has a crown–rump length of $11.8\ \text{cm}.$

Units of Mass

Mass is a measure of the amount of matter. Mass is measured on a balance or scale. The base unit of mass in the metric system is the **gram** (g). The SI unit of mass is the kilogram (kg). For example, when you read a nutritional label, such as the multivitamin shown in **Figure 1-11**, the mass of each vitamin and mineral in a tablet is listed. This nutritional label shows that one tablet contains $60\ \text{mg}$ of vitamin C.

Catherine Bausinger

Figure 1-11 The nutritional label for a multivitamin showing the mass of each vitamin in a tablet.

Since the symbol for the prefix *micro* is the Greek letter, μ, the microgram is sometimes abbreviated mcg in nutritional applications, avoiding confusion with the use of a Greek letter. For example, the nutritional label in Figure 1-11 shows that a tablet contains $25\ \text{mcg}$ of vitamin K.

$$1\ \mu\text{g} = 1\ \text{mcg} = 10^{-6}\ \text{g}$$

The common English units of mass are the pound (lb) and the ounce (oz) and the conversion between them is shown below:

$$1\ \text{pound}\ \ (\text{lb}) = 16\ \text{ounces}\ \ (\text{oz})$$

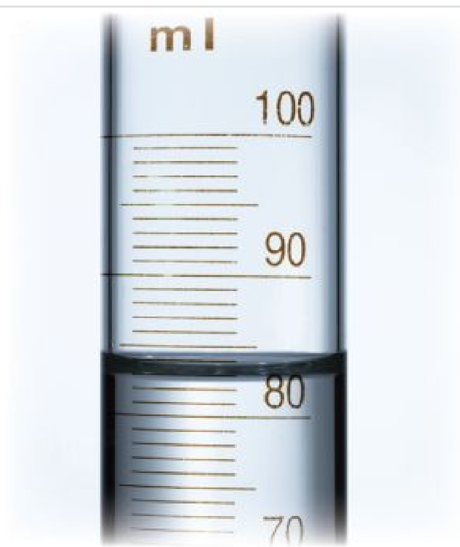
Units of Volume

Volume is a measure of three-dimensional space. One pint of blood (**Figure 1-12**) and $1\ \text{cc}$ of epinephrine (adrenaline) volume units used in the medical field.

Davies and Starr/The Image Bank/Getty Images

Figure 1-12 One pint of blood is a common volume measurements in the medical field.

In the laboratory, volume is typically measured with a graduated cylinder, shown in **Figure 1-13**. A liquid in a graduated cylinder forms a curved surface, known as a **meniscus**. A volume measurement should always be read from the *bottom* of the meniscus while viewing the meniscus at eye-level.



GIPhotoStock / Science Source

Figure 1-13 In the laboratory, volume is typically measured with a graduated cylinder. We read the volume at the bottom of the meniscus (curve).

Volume is a measurement derived from units of length. For example, a cube measuring 1 cm per side has a volume of $1\text{ cm}\times1\text{ cm}\times1\text{ cm}=1\text{ cm}^3$, as illustrated in one of the small cubes taken from the larger cube in Figure 1-14. The volume of this cube is read as "one centimeter cubed." Another term for this unit, common in medicine, is "one cubic centimeter," abbreviated 1 cc.

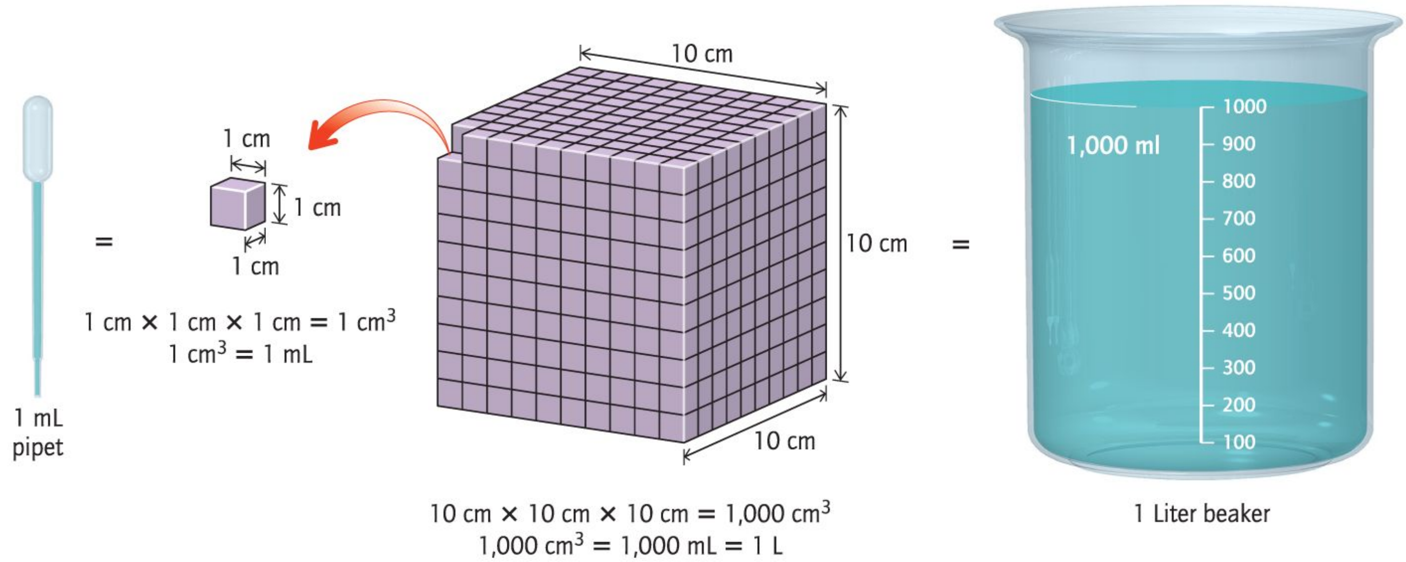


Figure 1-14 A small cube removed from the larger cube, measuring 1 cm along each side, has a volume of $1\text{ cm}^3=1\text{ mL}.$ The larger cube, measuring 10 cm along each side, has a volume of $1{,}000\text{ cm}^3=1{,}000\text{ mL}=1\text{ L}.$

The metric base unit of volume is the liter (L). One liter is the volume of a cube measuring 10 cm a side: $10\text{ cm}\times10\text{ cm}\times10\text{ cm}=1{,}000\text{ cm}^3=1\text{ L}$, as shown for the large cube in Figure 1-14. Therefore, $1\text{ L}=1{,}000\text{ cm}^3$; The large cube is the equivalent of a thousand of the smaller cubes. This gives us the important conversion between the milliliter and the cubic centimeter:

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

The common English units of volume are the gallon (gal), the quart (qt), and the pint (pt). The conversions between these English units is shown below:

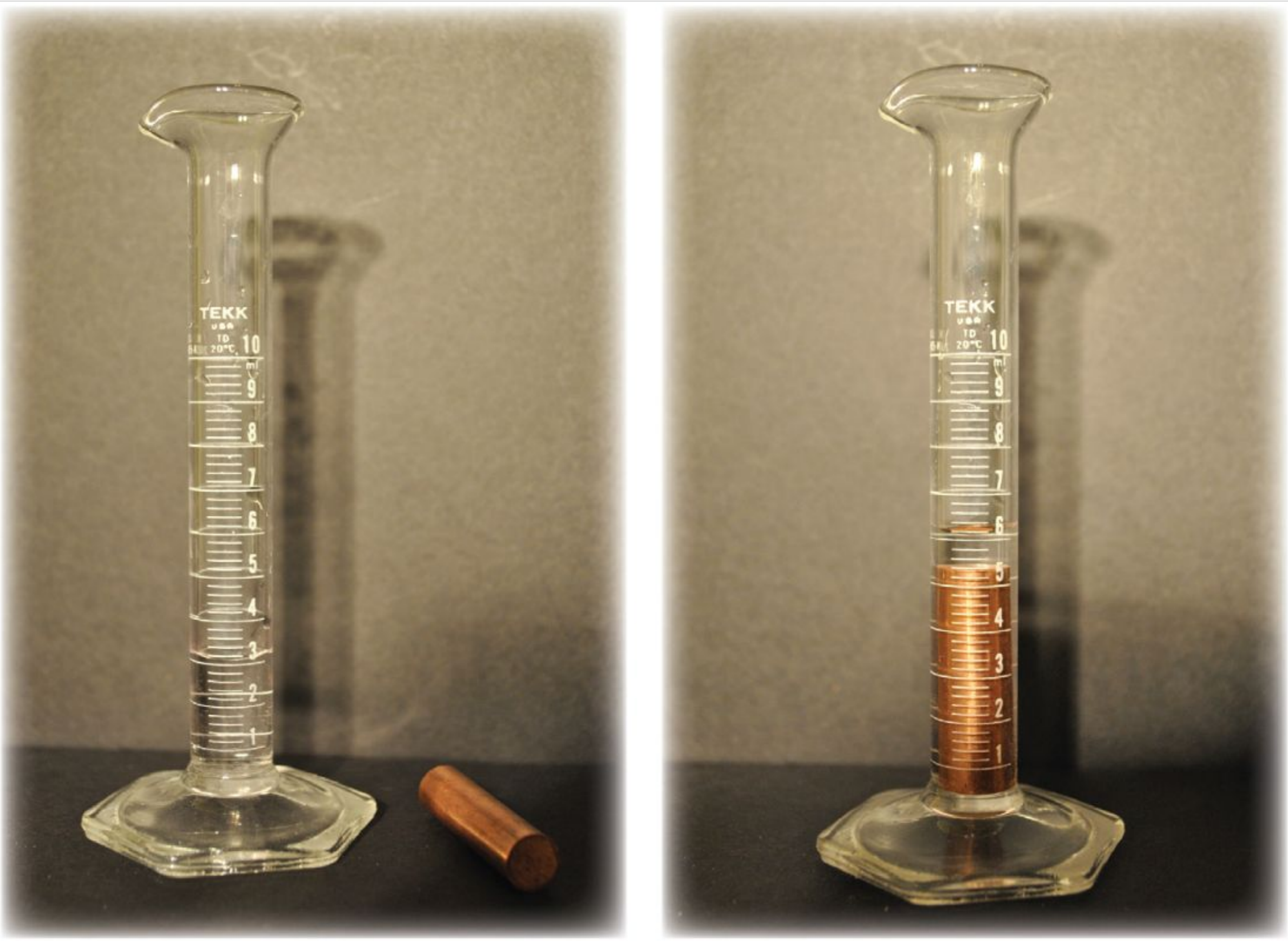
$$\begin{aligned}1\text{ gallon (gal)} &= 4\text{ quarts (qt)}\\ 1\text{ quart (qt)} &= 2\text{ pints (pt)}\end{aligned}$$

CORE CONCEPT: Volume is a unit of measurement derived from length where $1\text{ mL}=1\text{ cm}^3$ and $1\text{ L}=1{,}000\text{ cm}^3$.

CORE CONCEPT: The volume of an object can be measured by determining the volume of water that it displaces:
$V_{object} = V_{final} - V_{initial}$

The volume of a solid object can be calculated from its length or radius if it has a uniform shape, such as the cube shown in Figure 1-14. Alternatively, and if the object has an irregular shape, the volume of a solid can be determined by *displacement*. For example, to determine the volume of the cylinder of copper shown in **Figure 1-15a** by displacement, the copper cylinder is submerged in the graduated cylinder after it is filled with a measured volume of water $(V_{\text{initial}} = 3.0\ \text{mL})$. The copper cylinder causes the volume of water in the graduated cylinder to increase $(V_{\text{final}} = 6.1\ \text{mL})$, as shown in **Figure 1-15b**. The volume of the copper cylinder is equal to the amount of water that it displaces, calculated by subtracting the initial volume of water from the final volume of water:

$$V_{object} = V_{final} - V_{initial}$$



(a) (b)

Catherine Bausinger

Figure 1-15 Determining volume by displacement: (a) A copper sample and a graduated cylinder containing an initial volume of water, $V_{initial} = 3.0\ \text{mL}$; (b) The volume of the graduated cylinder, $V_{final} = 6.1\ \text{mL}$, after the copper sample was carefully added to the graduated cylinder.

WORKED EXERCISES

Metric and English Units of Length, Mass and Volume

1-5 Write the conversion between the gram and the microgram using the symbols for these units. Are these units of length, mass, or volume? Are these units from the English or the metric system?

Solution:

The gram (g) is the base unit of mass in the metric system and the microgram $(\mu\mathrm{g})$ is a prefixed metric unit of mass. From Table-1-1, we see that the symbol μ is the abbreviation for *micro*, and represents the multiplier 10^{-6}. Therefore, to write the conversion, we let 10^{-6} stand in for μ in $\mu\mathrm{g}$:$1\ \mu\mathrm{g} = 10^{-6}\ \mathrm{g}.$

1-6 Write the conversion between the meter and the kilometer using the symbols for these units. Are these units of length, mass, or volume? Are these units from the English or the metric system?

Solution:

The meter is the base unit for length and distance in the metric system and the kilometer (km) is a prefixed metric unit of distance. From Table 1-1, we see that the symbol *k* is the abbreviation for *kilo*, and represents the multiplier 10^3. Therefore, to write the conversion: $1\ \mathrm{km} = 10^3\ \mathrm{m}.$

1-7 When an irregular piece of gold was placed into a graduated cylinder containing $100.0\ \mathrm{mL}$ of water, the volume of water increased to $120.0\ \mathrm{mL}.$ What is the volume of the piece of gold in units of mL? What is the volume of the piece of gold in units of cm^3?

Solution:

The volume of a solid can be measured by displacement. The volume of the gold is equal to the difference in the volume of the water before and after the gold piece was submerged: $\mathrm{V_{gold}} = 120.0\ \mathrm{mL} - 100.0\ \mathrm{mL} = \mathbf{20.0\ mL}.$ To convert from units of mL to units of cm^3, we use the conversion: $1\ \mathrm{mL} = 1\ \mathrm{cm}^3.$ Therefore, $20.0\ \mathrm{mL} = \mathbf{20.0\ cm^3}.$

PRACTICE EXERCISES

7. Write the conversion between the mile and the foot. Are these units of distance, mass, or volume? Are these units from the English or the metric system?

8. Which of the following measurements represent English units and which represent metric units? Also, identify each as a unit of length, mass, or volume.
   a. $0.5\ \mu\mathrm{m}$
   b. $6\ \mathrm{gal}$
   c. $100\ \mathrm{kg}$
   d. $8\ \mathrm{cm}^3$
   e. $125\ \mathrm{lb}$
   f. $0.6\ \mathrm{mL}$

9. Which of the measurements below are diameters of objects visible to the naked eye?

a. 1 km
b. 1 m
c. 1 nm
d. 0.5 in.

10. What is the volume of an object, in cubic centimeters, if it is placed in $10. \text{ mL}$ of water and causes the volume of the water to increase to 11 mL?

11. What is the volume of a cube measuring 2 cm a side, in units of milliliters?

1.3 Significant Figures in Measurements and Calculations

Every measurement contains a degree of uncertainty that is due to the inherent limitations of the measuring device and our ability to read it. The degree of uncertainty in a measurement is expressed by the number of digits reported in the numerical value.

Significant Figures in a Measurement

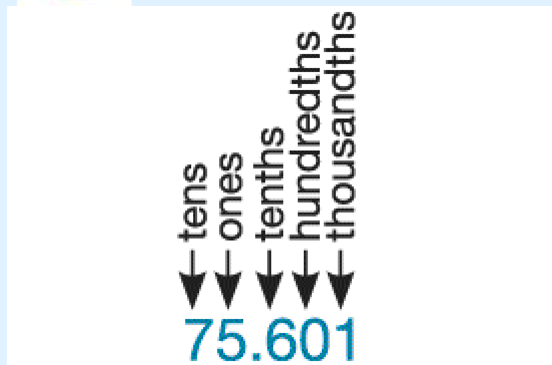
Even with the most sophisticated measuring device, a measurement can never be known exactly. The convention is to record *all* certain digits (those we are sure of) and estimate *one* uncertain digit. The last digit in a measurement is the uncertain digit. All the certain digits plus the one uncertain digit are known as **significant figures**.

To interpret a measurement, we can assume the last significant figure has been estimated. For example, a measurement reported as $5.62\ \text{cm}$ can be interpreted as a length somewhere between $5.61\text{–}5.63\ \text{cm}\ (5.62 \pm 0.01\ \text{cm})$, since the last significant figure, 2, is in the hundredths place.

If a less precise ruler had been used and that same length was reported as $5.6\ \text{cm},$ it would have been interpreted as somewhere between $5.5\text{–}5.7\ \text{cm}\ (5.6 \pm 0.1\ \text{cm}),$ since the last significant figure, 6, is in the tenths place. Since the second measurement has two significant figures, it has less certainty than the first measurement, which has three significant figures.

The type of measuring device used to take a measurement determines the number of significant figures that can be reported. For example, the blue solid shown in **Figure 1-16a**, weighed on the top-loading balance, displays a mass of $10.4\ \text{g},$ a value with *three* significant figures. The same sample weighed on the more precise analytical balance shown in **Figure 1-16b** displays a mass of $10.4977\ \text{g},$ a value with six significant figures. The last digit in both digital displays, which often fluctuates as you are taking the reading, is the uncertain digit. The more precise balance has the capacity to read more digits with certainty. The balance you choose to use depends on how much certainty you need for a given experiment.

MATH TIP:



CORE CONCEPT: Uncertainty in a measurement is conveyed through the number of significant figures reported: The last significant figure in a measurement is the uncertain and estimated digit.

(a) (b)

Figure 1-16 A blue solid shown on: (a) A top-loading balance reading to the tenths. (b) An analytical balance reading to the ten thousandths place (four places past the decimal).

Figure 1-17 Common laboratory glassware used for measuring volume: a syringe, a volumetric flask, a volumetric pipette, and a graduated cylinder.

Specially marked glassware is used to measure the volume of a liquid, such as the volumetric flask, volumetric pipet, graduated cylinder, and syringe shown in Figure 1-17. In the laboratory and in the clinic, it is good practice to record the correct number of significant figures when taking measurements. Do not estimate more than one digit; but also, do not omit any certain digits or the estimated digit. A common mistake is to drop a zero following a decimal, thinking it doesn't add value because it is zero. For example, if you measure a length to be $4.0\ \text{cm},$ don't drop the zero and simply report " $4\ \text{cm}.$ " A measurement of $4\ \text{cm}$ conveys less certainty than a measurement of $4.0\ \text{cm}.$ A measurement reported as $4\ \text{cm}$ (1 significant figure) suggests to someone interpreting the measurement that it is between $3–5\ \text{cm}\,(4 \pm 1);$ whereas a measurement of $4.0\ \text{cm}$ (2 significant figures) conveys the measurement is between $3.9–4.1\ \text{cm}\,(4.0 \pm 0.1),$ a measurement with more certainty.

CORE CONCEPT: Zeros that serve as placeholders in a number are not significant figures.

Zeros can be confusing because they are not always significant. Zeros are not significant when they serve as *placeholders* in a number. For numbers less than 1, these are the zeros between the decimal and the first nonzero digit, such as the three pink zeros in 0.00050, a number with two significant figures. For numbers greater than 1, these are the trailing zeros after a nonzero digit when there is no decimal, such as the four pink zeros in 70,000, a number with one significant figure. Guidelines for identifying and counting the significant figures in a number are summarized in [Table 1-2](). Examples are also provided.

Table 1-2 Guidelines for Identifying and Counting the Significant Figures in a Measurement

Digits that are Significant Figures	Example (in blue)	Number of Significant Figures
All nonzero digits	1.234	4
Zeros between nonzero digits (regardless of whether there is a decimal)	3.05 2006	3 4
Zeros to the right of a nonzero digit in a number containing a decimal	0.0400 20.00	3 4
A decimal placed after one or more zeros indicates the zeros are significant. Applies to numbers greater than 1	11,000. 11,000	5 2
All digits in the coefficient of a number expressed in scientific notation	2.30×10^4	3
Digits that are *not* Significant Figures (zeros are placeholders)	**Example (in pink)**	**No. of Significant Figures**
For numbers less than 1, the zeros between the decimal and the nonzero digit are not significant (they are placeholders).	0.00025 0.40	2 2
For numbers greater than 1, the zeros following the last nonzero digit (trailing zeros) are not significant (they are place holders).	6,000 340	1 2

To ensure that a decimal point is not overlooked in a number less than 1, a zero should be placed before the decimal point. This practice is routine in science and in the medical profession, especially when indicating dosages. Even an experienced nurse is more likely to miss the decimal point when the value is written as .32, for example, rather than 0.32, and mistaking it for 32, a value 100 times greater than 0.32.

Exact Numbers

Numbers obtained by accurate counting contain no uncertainty and are known as **exact numbers**. For example, if you carefully counted the number of students in your chemistry class and determined that there were 52 students, the number 52 would have no uncertainty. Since there is no uncertainty, there is no limit on the number of significant figures, and 52 is considered an exact number.

CORE CONCEPT: Exact numbers are not measurements. They include numbers obtained by accurate counting and defined conversions.

Defined quantities are also exact numbers, such as all the metric conversions in Table 1-1. Also, when multiplying or dividing a measured value by a whole number, like 2 or 3, such as when dividing a dose of medication, treat the whole number as an exact number.

WORKED EXERCISES

Significant Figures and Exact Numbers

1-8 For each of the following, indicate whether it is an exact number or a measured value. If it is a measured value, indicate the number of significant figures and underline the estimated digit.

a. $4.507\ \mathrm{cm}$

b. $0.00560\ \mathrm{g}$

c. $2.0 \times 10^5\ \mathrm{m}$

d. $53{,}000.\ \mathrm{seconds}$

e. $189\ \mathrm{students}$

Solution:

a. 4.507 cm. Measured value because it is a length. Four significant figures. All the digits are significant because nonzero digits are always significant and a zero between nonzero digits is significant. The last digit is the uncertain digit; it could be a 6, 7, or 8.

b. 0.00560 g. Measured value because it is a mass. Three significant figures. The zeros between the decimal and the first digit, 5, are not significant, they are placeholders. The zero after the last digit, 6, is significant and it is the uncertain digit.

c. 2.0 = 10^5 m. Measured value because it is a length. Two significant figures. All the digits in the coefficient of a number in scientific notation are significant. The last digit in the coefficient is uncertain.

d. 53,000. s. Measured value because it is time. Five significant figures. The decimal after the zeros indicates that the three zeros following the 3 are significant. The last of these zeros is uncertain.

e. Exact number because 189 students is a number obtained by counting.

1-9 What is the volume of the liquid in the graduated cylinder shown at right? Which digit(s) in your measurement are certain? Underline your estimated digit. How many significant figures does your measurement contain?

Solution:

$15.\underline{0}\ \mathrm{ML}.$ The 15 is certain, but the tenths place is uncertain. Three significant figures.

PRACTICE EXERCISES

12. For each of the following, indicate whether it is an exact number, or a measured value. If it is a measured value, indicate the number of significant figures and underline the estimated digit.
    a. $0.007\ \mathrm{m}$

b. 50 people

c. $23{,}000.$ seconds

d. 0.004050 mg

e. 3200 ft

13. The measurements below were made on three different balances:

i. 5.5 g

ii. 5.52 g

iii. 5.5093 g

a. Underline the estimated digit in each measurement.

b. A student takes a reading from a balance where the display fluctuates between $5.50\ \mathrm{g}$ and $5.53\ \mathrm{g}.$ Which of the three measurements above would be acceptable for the student to report?

c. How many significant figures does each measurement have?

Significant Figures in Calculations

Measurements are frequently used in calculations, as for example calculating the volume of a cube from measurements of length, as seen in the previous section. In a calculation involving measurements, it is incorrect to simply report all the digits displayed by the calculator. The correct number of significant figures to report in the final answer depends on the number of significant figures in each measurement and the mathematical operations used in the calculation: multiplication/division or addition/subtraction. A few simple rules are used to determine the correct number of significant figures to report in the answer when carrying out calculations using measurements.

Rules for Multiplication and Division

CORE CONCEPT: In calculations involving multiplication or division of measured values, round the final answer such that it has the same number of significant figures as the measurement with the fewest number of significant figures.

In calculations involving multiplication or division of measurements, the final answer cannot have more significant figures than the measurement with the *fewest* number of significant figures. For example, if you were to calculate the volume of a room by multiplying the length, width, and height of the room as shown below, the answer displayed on a TI-30XA calculator would show seven digits. If you reported all seven digits, you would be indicating that the volume of room was known with greater certainty than the measurements used to calculate the volume, which is impossible. A calculator cannot increase the certainty of a measurement!

$4.570\ \mathrm{m}$	$\times$	$3.7\ \mathrm{m}$	$\times$	$2.74\ \mathrm{m}$	$=$	$< 46.33066\ \mathrm{m}^3 >$	$=$	$46\ \mathrm{m}^3$
4 significant figures		2 significant figures		3 significant figures		7 significant figures		2 significant figures

We must round the answer to the correct number of significant figures. In this example, we determine that the measurement with the fewest number of significant figures is 3.7 (2 significant figures). Therefore, the final answer should also have two significant figures. We round 46.33066 to 46, reporting a final answer of $46\ \mathrm{m}^3.$

MATH TIP: Just as calculators can display too many digits, they drop zeros when they are not placeholders. Remember to add zeros back if they are significant figures.

In calculations that involve more than one mathematical step, do not round after each step; instead, carry the extra digits and round only the final answer to avoid rounding errors.

MATH TIP: When rounding a number, look at the digit immediately to the right of the digit you are rounding: if it is greater than or equal to 5, round up; if it is less than 5 leave it as is. Then drop all the digits to the right of the digit being rounded. Examples:

The number:	84.388
rounded to two significant figures:	84
rounded to three significant figures:	84.4
rounded to four significant figures:	84.39

Rules for Addition and Subtraction

CORE CONCEPT: In a calculation involving addition or subtraction of measured values, the answer cannot have more places to the right of the decimal than the measurement with the fewest places to the right of the decimal.

In calculations involving addition or subtraction of measurements, the final answer cannot have more places to the right of the decimal than the measurement with the fewest places to the right of the decimal. For example, the sum of the measurements shown below should be reported as $19.3\ \text{cm}$ because the measurement with the fewest places to the right of the decimal, $5.4\ \text{cm},$ has its last digit one place to the right of the decimal. Thus, round 19.294 to 19.3.

5.4
6.55
+ 7.344
<19.294> Calculator answer
19.3 cm Final answer after rounding

WORKED EXERCISE

Significant Figures in Calculations Involving Measured Values

1-10 Perform the following calculations by rounding the final answer to the correct number of significant figures:

a. $0.0022\ \text{cm} \times 58.88\ \text{cm} =$

b. $7.0\ \text{in.} + 8.55\ \text{in.} + 233\ \text{in.} =$

Solution:

a. Since the calculation involves multiplication, the calculated answer cannot have more significant figures than the measurement with the fewest significant figures. Since $0.0022\ \text{cm}$ has two significant figures and $58.88\ \text{cm}$ has four significant figures, the final answer should have two significant figures. The calculator shows a value of 0.129536, so we round up to $0.13\ \text{cm}^2$.

b. Since the calculation involves addition, the final answer cannot have more digits to the right of the decimal than the measurement with the fewest digits to the right of the decimal. The value 233 has no digits to the right of the decimal (the last digit is in the ones place). Therefore, the final answer must be rounded to the ones place. The calculator shows a sum of 248.45, which is rounded to 248 in.

$$\begin{array}{rl} 7.0 & \\ 8.45 & \\ +\ 233. & \\ \hline <\ 248.45\ > & \text{Calculator answer} \\ 248\ \text{in.} & \text{Final answer after rounding} \end{array}$$

PRACTICE EXERCISE

14. Perform the following calculations by rounding the final answer to the correct number of significant figures:

a. $56.50\ \text{m} \times 37.99\ \text{m} =$

b. $5.987\ \text{g} + 6.001\ \text{g} + 3.22\ \text{g} =$

c. $5.70\ \text{g} \times \dfrac{1.0\ \text{mL}}{1.8\ \text{g}} =$

1.4 Dimensional Analysis

CORE CONCEPT: Unit conversions—converting a measurement from one unit into another unit—are reliably solved using dimensional analysis.

Scientists and medical professionals routinely perform calculations. Many calculations, such as dosage calculations, are critical to the health and well-being of the patient. The most common type of calculation in science is a **unit conversion**, turning a measurement given in one unit into its equivalent in another unit. Unit conversions can be done between two metric units (the same system of measurement) or between a metric unit and an English unit (different systems of measurement). For example, converting your body weight from pounds to kilograms is an English-metric unit conversion. Unit conversions and many other calculations that you will be performing are reliably solved using a process known as dimensional analysis.

Using Dimensional Analysis in Metric-Metric Unit Conversions

Dimensional analysis is a process for solving mathematical problems that focuses on the *units*—the dimensions—in the problem, to direct how the calculation is set up. Using dimensional analysis can help avoid making many common calculation errors.

Metric Conversions with One Conversion Factor

The basic steps for solving a problem using dimensional analysis are given in the *Guidelines: Using Dimensional Analysis for a Unit Conversion Part I*. These guidelines illustrate the steps in a unit conversion between two metric units, one of which is the metric base unit and the other a prefixed metric unit. This is the simplest type of dimensional analysis problem because it requires only one conversion.

Guidelines: Using Dimensional Analysis for a Unit Conversion (Part I)

Conversions with *One* Conversion Factor

Example: *Convert* $0.500\ \text{g}$ *into mg.*

Step 1: Identify the *given* and *asked for* units in the problem. Identify the information that has been *given* and determine what you are being *asked for* by focusing on the *units*.	*Given*: $0.500\ \text{g}$ *Asked for*: mg
Step 2: Identify the conversion between the *given* units and *asked for* units. Determine if there is a well-known conversion, a mathematical equality, between the *given unit* and the *asked for* unit. Consult appropriate tables. Table 1-1 provides conversions between the common prefixed metric units and the base unit.	*In this example, we need a conversion that equates grams* (g) *and milligrams* (mg), *a metric-to-metric conversion. From Table 1-1:* $$1\ \text{mg} = 10^{-3}\ \text{g}$$
Step 3: Express the conversion as two *conversion factors*. Any conversion can also be expressed as a ratio, known as a **conversion factor**. One side of the equality becomes the numerator and the other side of the equality becomes the denominator. Then, invert the numerator and the denominator to obtain the other conversion factor.	*The conversion in step 2 can be expressed as two conversion factors:* $$\frac{10^{-3}\ \text{g}}{1\ \text{mg}} \text{ and } \frac{1\ \text{mg}}{10^{-3}\ \text{g}}$$
Step 4: Set up the calculation by writing the *given* from step 1 (both the numerical value and the *unit*) and *multiplying* by the conversion factor from step 3 that has the same unit as the *given* in the *denominator*. This will allow the *given* units to cancel, leaving only the *asked for* unit in the answer. If the asked for unit does not appear in the answer, the wrong conversion factor was selected. Select the other conversion factor and try again.	*For the example above, we write* $0.500\ g$, *and then multiply by the conversion factor which has grams in the denominator, the same unit as the given:* $$\underbrace{0.500\ \cancel{\text{g}}}_{\text{given}} \times \underbrace{\frac{1\ \text{mg}}{10^{-3}\ \cancel{\text{g}}}}_{\text{conversion factor}} = \underbrace{\_\_\_\_\ \text{mg}}_{\text{asked for}}$$ *Grams cancel in the given (a numerator) and the denominator of the conversion factor, leaving only milligrams, the asked for unit. This confirms that the calculation has been set up correctly before inputting any numerical values into the calculator.*
Step 5: Solve the numerical part of the problem and round to the correct number of significant figures. Multiply values that appear in the numerators and divide by values that appear in the denominator. Check the number of significant figures in the answer.	*For the problem above:*

$$0.500\ \cancel{g} \times \frac{1\ \text{mg}}{10^{-3}\ \cancel{g}} = 500.\ \text{mg}$$

We arrive at this answer by performing the following division operation $(0.500/10^{-3}) = 500.$ *Finally, we double check that we have recorded the correct number of significant figures.*

MATH TIP: The left side of the equation above can be interpreted as:

$$\frac{0.500\ \text{g} \times 1\ \text{mg}}{10^{-3}\ \text{g}}$$

WORKED EXERCISE

Dimensional Analysis with One Conversion Factor

1-11 The average healthy adult has $5\ \text{L}$ of blood in their circulatory system. What does this volume correspond to in milliliters?

Solution:

Step 1: We identify the given and asked for by focusing on the units.

Given: $\underline{5}\ \text{L}$ (1 significant figure)

Asked for: mL

Step 2: Recognize that both units are metric units of volume. One is a prefixed unit of volume, mL, and the other is the base unit of volume, L. Therefore, we need the conversion between the liter and the milliliter. To determine the conversion, we must determine the multiplier (the exponent) represented by the prefix, as described in section 1.2:

$$1\ \text{mL} = 10^{-3}\ \text{L}$$

Step 3: Express this conversion as two ratios:

$$\frac{1\ \text{mL}}{10^{-3}\ \text{L}} \quad \text{and} \quad \frac{10^{-3}\ \text{L}}{1\ \text{mL}}$$

Step 4: We set up the calculation by writing the given, $5\ \text{L},$ and multiplying by the first conversion factor in step 3 because it has liters in the denominator. Liters cancel because they appear in the numerator of the given and the denominator of the conversion factor, leaving only milliliters, the asked for unit. This indicates we have set up the calculation correctly. Then solve the numerical part of the calculation with the aid of a calculator:

$$5\ \cancel{L} \times \frac{1\ \text{mL}}{10^{-3}\ \cancel{L}} = 5{,}000\ \text{mL}$$

MATH TIP: The sequence of numbers and operations we input into our calculator is: $5 \div 1\ \text{EE} \pm 3 =$

PRACTICE EXERCISES

15. Convert $77{,}000\ \text{grams}$ into kilograms. Which unit is more convenient for this quantity, grams or kilograms?

16. Convert $561\ \text{mL}$ of water into liters.

17. A tumor has a diameter of $5.5\ \text{nm}.$ What is the diameter of the tumor in meters?

18. A tablet of regular strength Aleve contains $220\ \text{mg}$ of naproxen, a nonsteroidal anti-inflammatory drug available over-the-counter. How many grams of naproxen are in a tablet of Aleve?

Metric Conversions with Two Conversion Factors

CORE CONCEPT: Dimensional analysis allows us to multiply two or more conversion factors, setting them up in sequence such that units cancel, until only the asked for unit remains.

In many problems, a conversion that equates the *given* and the *asked for* units is not commonly known or available in tables. In fact, it is unnecessary to have a conversion between every possible set of units because the *given* can be multiplied by *two,* or more, conversion factors that are related to each other by another unit. The conversion factors are set up in sequence such that all units cancel, except the *asked for* unit, as described in the *Guidelines: Using Dimensional Analysis for a Unit Conversion Part II*.

The conversion between a prefixed metric unit and another prefixed metric unit is a common unit conversion that is readily solved using *two* conversion factors. The first conversion factor is between the *given* and the *base* unit, and the second conversion factor is between the *base* unit and the *asked for* unit. Thus, the two conversion factors are related to each other through a common unit, the *base* unit.

In the sections that follow, we describe how dimensional analysis can be used to perform metric-English unit conversions, density calculations, and dosage calculations, demonstrating the versatility of this method of solving problems.

Guidelines: Using Dimensional Analysis for a Unit Conversion (Part II)

Conversions with *Two or More* Conversion Factors

Example: Convert $78{,}000\ cm$ into km.

Step 1: Identify the *given* and *asked for* units in the problem. Identify the information that has been *given* and determine what is being *asked for* by focusing on the *units*.

Given: $78{,}000\ \text{cm}$

Asked for: km

Step 2: Identify the conversion between the *given* units and *asked for* units. If a direct conversion is not available, identify a conversion between the *given* unit and an *intermediate* unit and a second conversion between the *intermediate* unit and the *asked for* unit. For metric conversions with prefixed metric units, the *base* unit always serves as the intermediate unit.

Table 1-1 doesn't provide a direct conversion between the cm and the km, but there are conversions between both prefixed units and the base unit, the meter (m). From Table 1-1:

$1\ \text{cm} = 10^{-2}\ \text{m}$ and $1\ \text{km} = 10^{3}\ \text{m}$

Step 3: Express all conversions as their two possible *conversion factors*.

The conversions in step 2 can be expressed as the following pairs of conversion factors:

$$\frac{10^{-2}\ \text{m}}{1\ \text{cm}} \text{ and } \frac{1\ \text{cm}}{10^{-2}\ \text{m}}$$

$$\frac{1\ \text{km}}{10^{3}\ \text{m}} \text{ and } \frac{10^{3}\ \text{m}}{1\ \text{km}}$$

Step 4: Set up the calculation by writing the *given* from step 1 and multiplying by the conversion factor from step 3 that has the same unit as the *given* in the *denominator*. The *given units* in the numerator and the denominator of the conversion factor will cancel, leaving only the *intermediate* unit in the *numerator*. Continuing on the same line, multiply by the conversion factor in step 3 that has the same unit as the *intermediate* unit in the *denominator* and the asked for unit in the *numerator*. The *intermediate* units will cancel, leaving only the *asked for* unit. For some calculations more than two conversion factors may be required.

Write the given, $78{,}000\ cm$ *and multiply by the first conversion factor shown below because it has centimeters (the given unit) in the denominator. The given units cancel, leaving only the intermediate unit, meters, in the numerator. Then multiply by the second conversion factor shown below because it has meters (the intermediate unit) in the denominator and the asked for unit, km, in the numerator. The intermediate units, meters, cancel, leaving only the asked for unit, km:*

$$\underset{\text{given}}{78{,}000\ \cancel{\text{cm}}} \times \underset{\text{conversion factor 1}}{\frac{10^{-2}\ \cancel{\text{m}}}{1\ \cancel{\text{cm}}}} \times \underset{\text{conversion factor 2}}{\frac{1\ \text{km}}{10^{3}\ \cancel{\text{m}}}} =$$

Step 5: Solve the numerical part of the problem. Multiply all values that appear in the numerator and divide by all values that appear in the denominator.

MATH TIP: Remember when multiplying numbers in a denominator to input the open and closed parentheses ( ) in the calculator.

For the problem above:

$$\underset{\text{given}}{78{,}000\ \cancel{\text{cm}}} \times \underset{\text{conversion factor 1}}{\frac{10^{-2}\ \cancel{\text{m}}}{1\ \cancel{\text{cm}}}} \times \underset{\text{conversion factor 2}}{\frac{1\ \text{km}}{10^{3}\ \cancel{\text{m}}}} =$$

There are two significant figures in $78{,}000\ cm$ so the final answer should have two significant figures. The conversions themselves are exact numbers. Remember to include a zero before the decimal point so that it isn't overlooked.

WORKED EXERCISE

Dimensional Analysis with Two Conversion Factors

1-12 The average length of the femur bone in a fetus at 20 weeks gestation is 32 mm. What is this length in decimeters?

Solution:

Step 1: We identify the *given* and *asked for* quantities, paying particular attention to the units.

Given: $32\ \mathrm{mm}$

Asked for: dm

Step 2: From the units, we note that they are both prefixed metric units, which tells us that we will need two conversions and that the intermediate unit is the base unit, the meter. Therefore, we need the conversion from the millimeter to the meter, and the conversion from the meter to the decimeter: $\mathrm{mm} \rightarrow \mathrm{m}$ and $\mathrm{m} \rightarrow \mathrm{dm}$.

Step 3: We express the first conversion, $\mathrm{mm} \rightarrow \mathrm{m}$, as two conversion factors:

$$\frac{1\ \mathrm{mm}}{10^{-3}\ \mathrm{m}} \quad \text{and} \quad \frac{10^{-3}\ \mathrm{m}}{1\ \mathrm{mm}}$$

and also express the second conversion, $\mathrm{m} \rightarrow \mathrm{dm}$, as two conversion factors:

$$\frac{1\ \mathrm{dm}}{10^{-1}\ \mathrm{m}} \quad \text{and} \quad \frac{10^{-1}\ \mathrm{m}}{1\ \mathrm{dm}}$$

Steps 4 and 5: Using dimensional analysis, set up the calculation by first writing the *given*: $32\ \mathrm{mm}$. Then, multiply by the conversion factor that has mm in the denominator and meters (the intermediate unit) in the numerator, so that mm cancel. We now have units of *meters* remaining in the numerator. This is our prompt that the next step is to multiply by the conversion factor that has meters in the denominator and decimeters in the numerator. *Meters* cancel leaving only the *asked for* unit, decimeters. This confirms we have set up the calculation correctly. We then solve the numerical part of our problem and report our answer to two significant figures:

$$32\ \cancel{\mathrm{mm}} \times \frac{10^{-3}\ \cancel{\mathrm{m}}}{1\ \cancel{\mathrm{mm}}} \times \frac{1\ \mathrm{dm}}{10^{-1}\ \cancel{\mathrm{m}}} = 0.32\ \mathrm{dm}$$

WORKED EXERCISE

Dimensional Analysis with Multiple Conversion Factors

1-13 Which of the following is the larger mass: $3.00\ \mathrm{mg}$ or $325\ \mu\mathrm{g}$?

Solution:

Step 1: In this question, we cannot simply compare the size of the numerical values because they have different units. To compare measurements, they must have the same units. Thus, we can either convert $3.00\ \mathrm{mg}$ to micrograms or convert $325\ \mu\mathrm{g}$ to milligrams. In this solution we describe the latter.

Given: $325\ \mu\mathrm{g}$

Asked for: mg

Step 2: Both the *given* and *asked for* units are prefixed metric units, so we need one conversion from the microgram to the gram, and one conversion from the gram to the milligram where the base unit, the gram, is the intermediate unit: $\mu\mathrm{g} \rightarrow \mathrm{g}$ and $\mathrm{g} \rightarrow \mathrm{mg}$.

Step 3: We express the first conversion, $\mu\mathrm{g} \rightarrow \mathrm{g}$, as two conversion factors:

$$\frac{1\ \mu\mathrm{g}}{10^{-6}\ \mathrm{g}} \quad \text{and} \quad \frac{10^{-6}\ \mathrm{g}}{1\ \mu\mathrm{g}}$$

and then express the second conversion, $\mathrm{g} \rightarrow \mathrm{mg}$, as two conversion factors:

$$\frac{1\ \mathrm{mg}}{10^{-3}\ \mathrm{g}} \quad \text{and} \quad \frac{10^{-3}\ \mathrm{g}}{1\ \mathrm{mg}}$$

Steps 4 and 5: Using dimensional analysis, set up the calculation by first writing the *given*: $325\ \mu\mathrm{g}$. Then, multiply by the conversion factor that has $\mu\mathrm{g}$ in the denominator and grams (the intermediate unit) in the numerator, so that $\mu\mathrm{g}$ cancels. We see that we have units of *grams* remaining in the numerator. This is our prompt that the next step is to multiply by the conversion factor that has grams in the denominator and milligrams in the numerator, so that *grams* cancel, leaving only the *asked for* unit, milligrams. This tells us we have set up the calculation correctly. We then solve the numerical part of our problem and report our answer to three significant figures:

$$325\ \cancel{\mu g} \times \frac{10^{-6}\ \cancel{g}}{1\ \cancel{\mu g}} \times \frac{1\ mg}{10^{-3}\ \cancel{g}} = 0.325\ mg$$

Step 6: Now we can compare this value, $0.325\ mg$, to the other given, $3.00\ mg$, to determine that $0.325\ mg (325\ \mu g)$ is greater than $3.00\ mg$.

PRACTICE EXERCISES

19. A cancer patient has a tumor $150.\ \mu m$ in diameter. What is the size of the tumor in millimeters? Use dimensional analysis to solve the problem.

20. The deciliter is a common unit of volume in medicine. If a patient receives $25\ dL$ of blood, how many milliliters of blood has the patient received? Use dimensional analysis to solve the problem.

21. Which of the following is a longer distance to run: $1.1\ km$ or $520{,}000\ cm$?

English-Metric Conversions

CORE CONCEPT: When the given unit and asked for unit are from different systems of measurement, an English-metric conversion is identified and used as one of the conversions using dimensional analysis.

In the previous section, we saw examples of unit conversions in which both units were from the metric system. When the *given* and *asked for* units are from different systems, an English-metric conversion is identified and used as one of the conversions when setting up a calculation using dimensional analysis. The common English-metric conversions for length, mass, and volume are given in Table 1-3a. A summary of the common English-English conversions is provided in Table 1-3b.

Table 1-3 English-Metric and English-English Conversions for Length, Mass, and Volume

Length	Mass	Volume
a. Metric-English and English-Metric Conversions		
1 meter (m) = 39.37 inches (in.) 1.09 yard (yd)	1 kilogram (kg) = 2.205 pounds (lb)	
1 kilometer (km) = 0.62 mile (mi)	1 ounce (oz) = 28.35 grams (g)	
1 inch (in.) = 2.54 centimeters (cm)		1 liter (L) = 1.06 quarts (qt)
1 foot (ft) = 30.48 centimeters (cm)		1 gallon (gal) = 3.79 liters (L)
b. English-English Conversions		
1 foot (ft) = 12 inches (in.)	1 pound (lb) = 16 ounces (oz)	1 gallon (gal) = 4 quarts (qt)
1 yard (yd) = 3 feet (ft)		1 quart (qt) = 2 pints (pt)

$1 \text{ mile (mi)} = 5{,}280 \text{ feet (ft)}$		

When a direct conversion between the *given* and *asked for* units is not available, two or more conversion factors will be required, such as additional metric-metric or English-English conversions.

WORKED EXERCISE

English-Metric Conversions

1-14 Suppose you weighed your patient and recorded her weight at $125 \text{ lb}.$ What is your patient's weight in kilograms?

Solution:

Step 1: Identify the *given* and *asked for* paying particular attention to the units.

Given: 125 lb (3 significant figures)

Asked for: kg

Step 2: From the *given* and *asked for* units, we recognize that pounds (lb) is an English unit of mass and that kilograms (kg) is a metric unit of mass. Thus, we need an English to metric mass conversion. From the second column of Table 1-3a we find that there is a direct conversion between pounds and kilograms, so only one conversion is needed:

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

Step 3: We express this conversion as the two conversion factors below (invert one conversion factor to get the other conversion factor):

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

Step 4: Using dimensional analysis, we set up the calculation by writing the *given*, $125 \text{ lb},$ and then multiply by the conversion factor that has pounds in the denominator. Pounds cancel, leaving kilograms in the numerator, the *asked for* unit.

$$125 \cancel{\text{lb}} \times \frac{1 \text{ kg}}{2.205 \cancel{\text{lb}}} = \_\_\_ \text{ kg}$$

Step 5: Using a calculator, we solve the numerical part of the calculation $(125/2.205 = 56.7),$ remembering to round to three significant figures:

$$125 \cancel{\text{lb}} \times \frac{1 \text{ kg}}{2.205 \cancel{\text{lb}}} = 56.7 \text{ kg (3 significant figures)}$$

PRACTICE EXERCISES

22. You are in Europe and you need to fill your empty 15–gallon gas tank. How many liters of gasoline do you need to purchase?

23. A premature baby weighs $906 \text{ g}.$ What is the weight of the baby in pounds?

24. You check the inventory in the blood bank where you work and report that there are 285 pints of Type O blood. Using only Table 1-3, how many liters of Type O blood does the blood bank have? Show your work.

Density Calculations

Density is a physical property of a substance (solid, liquid, or gas). For example, bone density is used to determine if a patient has osteoporosis (see *Concepts in Context: Osteoporosis and Measurement of Bone Density*). The **density** (d) of a substance is calculated from the mass (m) and volume (V) of a sample of the substance, according to the equation:

$$d = \frac{m}{V}$$

where the italicized letter *m* stands for mass, not to be confused with the abbreviation for the meter, m, which is not italicized.

Since density is calculated from mass *divided* by a volume, it has units of $\text{g/mL},$ read "grams per milliliter," or $\text{g/cm}^3.$ The density of water, for example, is $1.00 \text{ g/mL}.$ The density of a substance is independent of the amount of the substance. In other words, the density of a drop of water is the same as a bathtub full of water. The density of some common substances is shown in **Table 1-4**.

Table 1-4 Density of Substances at 25 °C

Substance Name	Density (g/cm^3)
Ethanol	0.79

Water	1.00
Iron	7.87
Gold	19.32

CORE CONCEPT: Density is defined as the mass divided by the volume of a sample of material. Density can be used as a conversion factor to calculate the mass or volume of a substance when given the density and either the volume or mass of the sample.

Density (*d*) *is* a ratio, and therefore, can be used as a conversion factor in calculations to determine the *mass* of a substance when the volume (*V*) and the density are given, or to determine the *volume* of a substance when the mass (*m*) and the density are given. This type of calculation requires only one conversion factor: density or the inverted form of density, as shown in the Worked Exercises on the next page.

Specific Gravity

Specific gravity, a measurement closely related to density, is used in medicine and veterinary medicine to rapidly screen for conditions related to kidney function. Specific gravity is defined as the ratio of the density of a substance to the density of water at $4\ ^\circ\mathrm{C}$:

$$\text{specific gravity} = \frac{\text{density of substance } \left(\frac{\mathrm{g}}{\mathrm{mL}}\right)}{\text{density of water } \left(\frac{\mathrm{g}}{\mathrm{mL}}\right)}$$

Since both the numerator and the denominator have the same units, specific gravity is unitless. Moreover, since the density of water is $1.00\ \mathrm{g/mL},$ the specific gravity of a substance has the same numerical value as its density but absent the units.

CORE CONCEPT: Specific gravity is a unitless measurement defined as the density of a substance divided by the density of water $(1.00\ \mathrm{g/mL}).$

The specific gravity of urine in a healthy individual ranges from 1.002 to 1.030. Values outside the normal range can be an indicator of diabetes, kidney failure, or kidney infection.

WORKED EXERCISES

Density and Specific Gravity Calculations

1-15 Calculate the density of a sample with a mass of $0.90\ \mathrm{g}$ and a volume of $1.2\ \mathrm{mL}.$

Solution:

To calculate density, we need the definition of density: $d = m/V,$ and we need to be given both the mass *and* the volume of the sample. Show your work by first writing the density equation and then substituting the given values into the density equation, including units:

$$d = \frac{m}{V} = \frac{0.90\ \mathrm{g}}{1.2\ \mathrm{mL}} = 0.75\ \mathrm{g/mL}\ \ (2\ \text{significant figures})$$

In this calculation, no units cancel, so the final answer has grams in the numerator and milliliters in the denominator $(\mathrm{g/mL}),$ the typical units of density. We use a calculator to divide the numerical values.

1-16 Gold has a density of $19.32\ \mathrm{g/cm^3}.$ Calculate the mass, in grams, of a $5.0\ \mathrm{cm^3}$ block of gold.

Solution:

Step 1: We identify the *given* and *asked for* quantities, paying particular attention to the units. We see that the volume and the density have been given.

Given: $5.0\ \mathrm{cm^3}$ and the density of gold: $\dfrac{19.32\ \mathrm{g}}{1\ \mathrm{cm^3}}$

Asked for: g

Steps 2 and 3: Density is already expressed as a conversion factor. To obtain the other conversion factor, invert the numerator and denominator:

$$\frac{19.32\ \text{g}}{1\ \text{cm}^3} \quad \text{and} \quad \frac{1\ \text{cm}^3}{19.32\ \text{g}}$$

density *inverted density*

Steps 4 and 5: We set up the calculation by writing the given, $5.0\ \text{cm}^3$. Then multiply by the conversion factor that has volume in the denominator so that volume cancels, leaving only grams in the numerator, the asked for unit:

$$5.0\ \cancel{\text{cm}^3} \times \frac{19.32\ \text{g}}{1\ \cancel{\text{cm}^3}} = 97\ \text{g of gold (2 significant figures)}$$

We round the final answer to two significant figures because the volume, $5.0\ \text{cm}^3$, has the fewest number of significant figures, two significant figures.

1-17 A patient's urine sample is found to have a specific gravity of 1.040. What is the density of this patient's urine? Is the specific gravity of the patient's urine within the normal range?

Solution:

The density of the patient's urine is $1.040\ \text{g/mL}$—specific gravity and density have the name numerical value. The normal range for the specific gravity of urine is 1.002–1.030, so at 1.040 the patient has a specific gravity above the normal range.

PRACTICE EXERCISES

25. What is the density of a liquid with a mass of $5.5\ \text{g}$ and a volume of $5.0\ \text{mL}$?
26. A piece of unknown metal was found to have a mass of $71.1\ \text{g}$ and a volume of $9.0\ \text{cm}^3$. Using Table 1-4, determine if this metal is gold or iron.
27. An object has a mass of $10.5\ \text{g}$. When it is submerged in a graduated cylinder containing $82.5\ \text{mL}$ of water, the volume of water increases to $95.0\ \text{mL}$. What is the density of the object?
28. Using Table 1-4, calculate the volume, in cm^3, of a $22\ \text{g}$ block of gold.
29. What is the mass, in grams, of $25\ \text{mL}$ of water?
30. Would you expect a woman with osteoporosis to have a bone density greater than or less than normal? Explain.
31. Ice floats in liquid water. Does this mean that ice has a density *greater than* or *less than* liquid water?

Dosage Calculations

Health care professionals who administer medications to patients learn that administering the correct dosage is critical. It is referred to as one of the *Five Rights of Medication Administration*:

•Correct Medication •Correct Patient •Correct Dosage •Correct Route •Correct time

It is essential that nurses and other health care professionals who administer medications know how to calculate the correct amount of medication to give to a patient. Sometimes the dosage prescribed for a patient is given per kilogram of body weight, especially for medications administered to children. For example, consider the order given below:

"$8.0\ \text{mg}$ of tetracycline per kilogram body weight q.d."

A mass or volume of medicine given per weight of the patient can be expressed as a conversion factor. It is the *mass* of the *medication* per *mass* of the patient's body *weight*. For the above example, the conversion factor is:

$$\frac{8\ \text{mg}}{1\ \text{kg}}$$

To calculate the correct amount of tetracycline to administer to the patient, we first obtain the patient's weight. If the patient's weight has been reported in pounds, we will need two conversion factors: the English-metric conversion for mass, from Table 1-3, and the dosage, expressed as the ratio above.

MATH TIP: The word **per** means "for every" and refers to a ratio, where per is represented by the division operation (/). For example, 60 miles per hour (mph) can be written $60\ \text{mi}/1\ \text{hr}$.

CORE CONCEPT: When dosage is given as a mass of medication per mass of body weight, it can be expressed as a ratio, which can be used as a conversion factor in dosage calculations using dimensional analysis.

Some common abbreviations used to indicate how often to administer a medication include ***q.d.*** and ***b.i.d.***, derived from the Latin words meaning administered "once daily" and "twice daily," respectively. If the medicine is prescribed *b.i.d*, divide your final answer by two to determine the amount of medicine to give the patient every $12\ \text{hours.}$

WORKED EXERCISE

Dosage Calculations

1-18 Tetracycline elixir, an antibiotic, is ordered at a dosage of $8.0\ \text{mg}$ per kilogram of body weight *q.d.* for a child weighing $52\ \text{lb.}$ How many milligrams of tetracycline elixir should be given to the child daily?

Solution:

Step 1: We identify the *given* and *asked for* quantities, paying particular attention to the units. We turn the dosage per kilogram body weight into a ratio by noting that *per* means divided by.

Given: $52.0\ \text{lb}$ and $\dfrac{8.00\ \text{mg}}{1\ \text{kg}}$

Asked for: mg a day

Step 2: Since the dosage is given in terms of the weight of the patient, kilograms, but the patient's weight is given in pounds, an English-to-metric conversion is also needed. From Table 1-3, second column: $1\ \text{kg} = 2.205\ \text{lb.}$

Step 3: We express the English-to-metric conversion as two conversion factors:

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

And note the dosage per kg, a conversion factor, from step 1:

$$\frac{8.00\ \text{mg}}{1\ \text{kg}}$$

Steps 4 and 5: We set up the calculation by writing the given weight (not the given conversion). Next, we multiply by the English-to-metric conversion factor that has pounds in the denominator. Pounds cancel, leaving kilograms in the numerator. Then, we multiply by the dosage conversion factor so kilograms cancel, leaving only mg in the numerator, the asked for unit and the amount of medication to give to the patient daily.

$$52.0\ \cancel{\text{lb}} \times \frac{1\ \cancel{\text{kg}}}{2.205\ \cancel{\text{lb}}} \times \frac{8.00\ \text{mg}}{1\ \cancel{\text{kg}}} = 189\ \text{mg}\ \text{(3 significant figures)}$$

We then multiply by all the numerical values in the numerator and divide by all the numerical values in the denominator to obtain a final answer of $189\ \text{mg}$ to be given once daily.

Do not be confused by the fact that there are two *mass* units in a dosage calculation. The mass of the patient is different from the mass of the medication; they should not be set up to cancel.

WORKED EXERCISE

Advanced Dosage Calculation

1-19 A doctor orders $500\ \text{mg}$ of Rocephin be given to a $30\ \text{lb}$ toddler every $8\ \text{hours.}$ The label on the medicine shows the recommended dosage is $75\text{–}150\ \text{mg}$ per kilogram body weight per day. Is the doctor's order within the recommended range for this medication?

Solution:

Step 1: We need to determine how much medication should be given to the toddler based on the recommended dosage and then compare it to what the doctor ordered. The *given* is the toddler's weight, $30\ \text{lb.}$ We consider both the low end dosage, $75\ \text{mg}$ per kg body weight and the high end dosage, $150\ \text{mg}$ per kg body weight as the two conversion factors that will give us the low and high end of dosages.

Given: $30\ \text{lb},\ \dfrac{75\ \text{mg}}{1\ \text{kg}}$ per day, and $\dfrac{150\ \text{mg}}{1\ \text{kg}}$ per day

Asked for: mg every $8\ \text{hours}$

Step 2: Since the dosage is given in terms of the weight of the patient in kilograms, but the patient's weight is given in pounds, an English-to-metric mass conversion is also needed. From Table 1-3, second column: $1\ \text{kg} = 2.205\ \text{lb.}$

Step 3: We convert the English-to-metric conversion into two conversion factors:

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

Steps 4 and 5: We set up the calculation by writing the given weight. Next, we multiply by the English-to-metric conversion factor that has pounds in the denominator. Pounds cancel, leaving kilograms in the numerator. Then, we multiply by the low end dosage conversion factor so kilograms cancel, leaving only mg in the numerator, the asked for unit and the *lowest* amount of medication to give to the patient *daily*.

$$30.\cancel{lb} \times \frac{1\ \cancel{kg}}{2.205\ \cancel{lb}} \times \frac{75\ mg}{1\ \cancel{kg}} = 1020\ mg\ \textit{per day}, \text{ which when divided into three equal doses}$$
$$= 340\ mg\ \textit{every 8 hours}$$

To look at the high end of the recommended dose, we do the same calculation but use the high end conversion factor:

$$30\ \cancel{lb} \times \frac{1\ \cancel{kg}}{2.205\ \cancel{lb}} \times \frac{150\ mg}{1\ \cancel{kg}} = 2041\ mg\ \textit{per day}, \text{ which when divided into three equal doses}$$
$$= 680\ mg\ \textit{every 8 hours}$$

Thus, a 30 lb child should receive between 340 mg and 680 mg every $8\ \text{hours}$. The doctor's order of 500 mg every $8\ \text{hours}$ is within the recommended range.

PRACTICE EXERCISES

32. Quinidine is an antiarrhythmic agent. It is prescribed for an adult patient weighing $110\ \text{lb}$ at a dosage of $25.0\ \text{mg}$ per kilogram of body weight *b.i.d*.
    a. Express the dosage as a conversion factor.
    b. How often should the medication be given?
    c. How much quinidine, in milligrams, should be given to the patient at each administration?

33. Ampicillin, an antibiotic, is prescribed for a child weighing $63.0\ \text{lb}$ at a dosage of $20.0\ \text{mg}$ per kilogram of body weight every day, in four equally divided doses throughout the day.
    a. How often should the medication be given?
    b. How many milligrams should be given at every administration?

34. The typical dose of Ivermectin, a treatment for the prevention of heartworm in dogs, is $6.00\ \text{micrograms}$ per kilogram body weight once a month. For dogs that carry the MDRI gene, 100. micrograms per kilogram body weight can be fatal. If you have a $95.0\ \text{lb}$ golden retriever, how many milligrams of Ivermectin should the dog be given every month for the prevention of heartworm? What would be a toxic dose, if the dog has the MDRI gene?

1.5 Temperature Scales and Conversions

In section 1.1 you learned that temperature is a measure of the average kinetic energy of a substance and that particles with a greater kinetic energy have a higher temperature than particles with less kinetic energy. In this section, we describe the temperature scales and how to convert between the different temperature scales.

The Temperature Scales

Phanie/Science Source

Figure 1-18 An infrared thermometer can be used to measure body temperature.

Temperature is measured with a thermometer, such as the infrared thermometer shown in **Figure 1-18**. Temperature is reported in one of three temperature scales: Fahrenheit $(^{\circ}\mathrm{F})$, Celsius $(^{\circ}\mathrm{C})$, or Kelvin (K).

The SI unit of temperature is the *kelvin (K)*. The **Kelvin scale** is an *absolute* temperature scale where 0 K is the temperature when particles have no kinetic energy, the lowest temperature theoretically possible. A degree sign $(^{\circ})$ is not used in the Kelvin scale because the units are kelvins and not degrees. The Kelvin scale is used in the sciences when an absolute temperature is required in a calculation, as for example, determining the pressure of a gas from the volume and temperature of the gas. We will perform calculations using kelvins in chapter 6 when we study gases. On the Kelvin scale, all temperatures have positive values.

CORE CONCEPT: Three temperature scales—Fahrenheit $(^{\circ}\mathrm{F})$, Celsius $(^{\circ}\mathrm{C})$, and Kelvin (K)—are used to report temperature. The Kelvin scale is an absolute temperature scale.

The Celsius and Fahrenheit scales are relative temperature scales where the units are degrees Celsius $(^\circ\mathrm{C})$ and degrees Fahrenheit $(^\circ\mathrm{F})$, respectively. The **Fahrenheit scale** $(^\circ\mathbf{F})$, is limited to certain applications in the United States, such as reporting the weather. The **Celsius scale** $(^\circ\mathbf{C})$ is the most commonly used temperature scale in the world. It is used universally in the sciences, medicine, and for everyday applications outside the United States.

Figure 1-19 compares the freezing point and the boiling point of water on all three temperature scales. Normal body temperature is also indicated on each scale: $98.6\ ^\circ\mathrm{F},\ 37\ ^\circ\mathrm{C},$ and $310\ \mathrm{K}.$ On the Celsius scale, the freezing point of water is $0\ ^\circ\mathrm{C}$ and the boiling point of water is $100\ ^\circ\mathrm{C},$ with 100 degree increments in between. On the Kelvin scale, the freezing point of water is $273\ \mathrm{K}$ and the boiling point of water is $373\ \mathrm{K},$ with 100 kelvins in between. On the Fahrenheit scale, the freezing point of water is $32\ ^\circ\mathrm{F}$ and the boiling point of water is $212\ ^\circ\mathrm{F},$ with 180 degree increments in between. Therefore, one degree Celsius and one kelvin are equivalent to 1.8 degrees Fahrenheit. For example, an increase from $0\ ^\circ\mathrm{C}$ to $10\ ^\circ\mathrm{C}$ (10 degrees) is equivalent to a change from $273\ \mathrm{K}$ to $283\ \mathrm{K}$ (10 kelvins). In contrast, this temperature change corresponds to an increase from $32\ ^\circ\mathrm{F}$ to $50\ ^\circ\mathrm{F},$ an 18 degree increase.

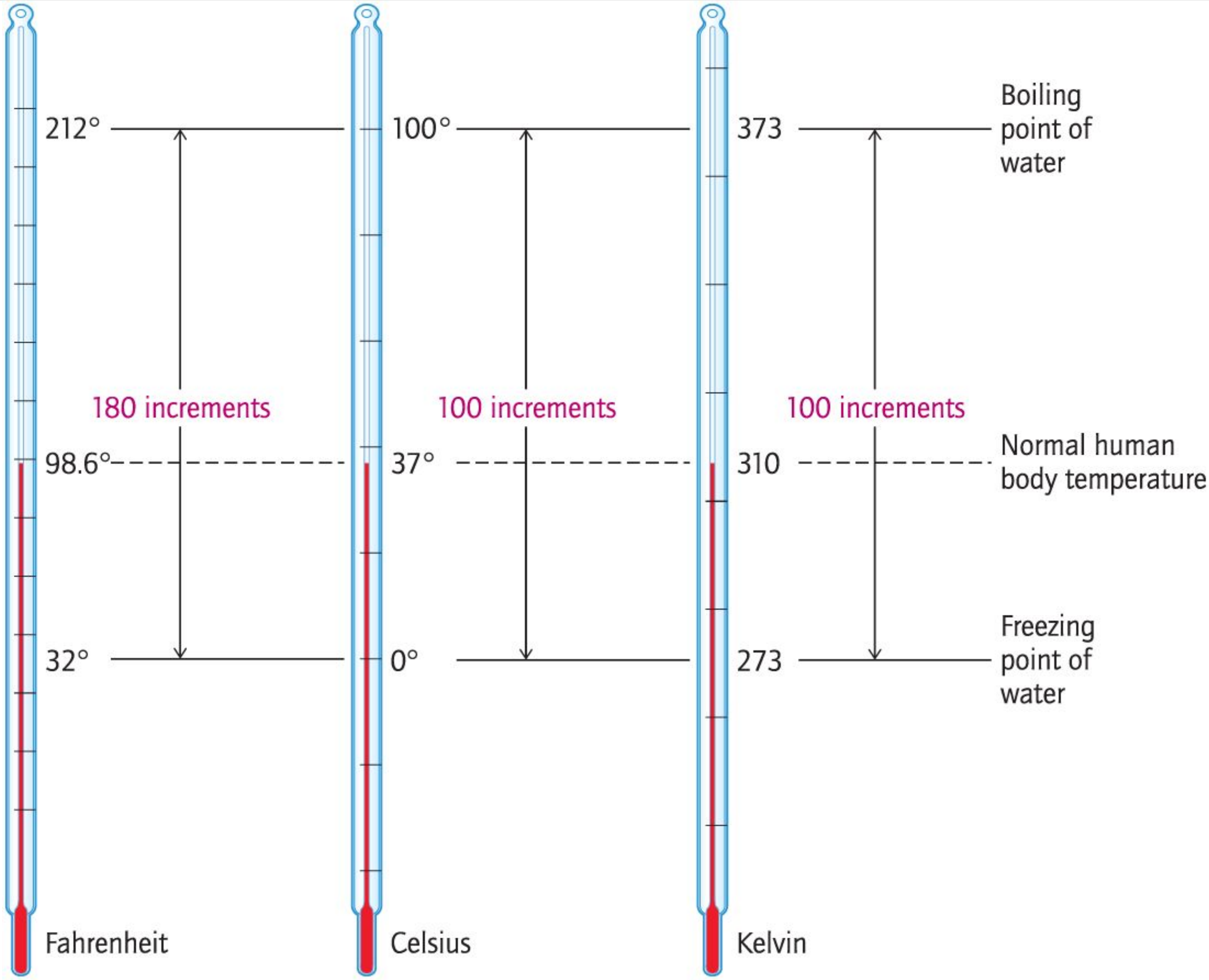


Figure 1-19 A comparison of the Fahrenheit, Celsius, and Kelvin temperature scales, showing the freezing point and boiling point of water on each scale. A comparison of normal human body temperature is also shown.

The human body must maintain its core temperature around $37\ ^{\circ}\mathrm{C}$. **Hypothermia** occurs when a person's core body temperature falls below $35\ ^{\circ}\mathrm{C}\ (95\ ^{\circ}\mathrm{F})$. Hypothermia is typically caused by prolonged exposure to cold weather or immersion in cold water for an extended period of time. At the other extreme, heat stroke occurs when the core body temperature rises above $41\ ^{\circ}\mathrm{C}\ (105\ ^{\circ}\mathrm{F})$.

Temperature Conversions

Use one or both of the equations given in Table 1-5 to convert a temperature from one temperature scale to another temperature scale. We cannot use dimensional analysis for this type of calculation because the scales are offset from one other.

Table 1-5 Equations for Interconverting between the Temperature Scales

$T_{^{\circ}\mathrm{F}} \rightarrow T_{^{\circ}\mathrm{C}}$	$T_{\mathrm{K}} \rightarrow T_{^{\circ}\mathrm{C}}$

Converting from $T_{^\circ F}$ to $T_{^\circ C}$	Converting from T_{K} to $T_{^\circ C}$
$T_{^\circ C} = \dfrac{(T_{^\circ F} - 32)}{1.8}$	$T_{^\circ C} = T_{K} - 273.15$

CORE CONCEPT: Two equations are used to interconvert between the temperatures scales. One equation is used for interconverting between $^\circ C$ and $^\circ F$ and another equation is used for interconverting between $^\circ C$ and K.

To perform a temperature conversion, the first step is to determine the *given* and the *asked for* temperature units. To interconvert between degrees Fahrenheit and Celsius, use the equation in the first column of Table 1-5; to interconvert between degrees Celsius and kelvins, use the equation in the second column of Table 1-5. To interconvert between degrees Fahrenheit and kelvins, first convert from degrees Fahrenheit to degrees Celsius and then from degrees Celsius to kelvins; thus, two sequential temperature conversions: $T_{^\circ F} \leftrightarrow T_{^\circ C} \leftrightarrow T_{K}$

For example, if the temperature is *given* in $^\circ C$ and you are *asked for* the temperature in $^\circ F$, the equation in the first column of Table 1-5 is required. However, the equation must first be algebraically rearranged and solved for $T_{^\circ F}$.

MATH TIP: To solve the temperature equation in the first column for $T_{^\circ F}$, rearrange the equation algebraically: $T_{^\circ F} = (1.8 \times T_{^\circ C}) + 32.0$

If instead the temperature is given in $^\circ C$, and you are asked for the temperature in K, the equation in the second column of Table 1-5 is required. To solve this equation for the temperature in kelvins, T_{K}, only one simple algebraic step is necessary.

MATH TIP: To solve the temperature equation in the second column for temperature in kelvins, K, add 273.15 to both sides and cancel like terms: $T_{K} = T_{^\circ C} + 273.15$

MATH TIP: Remember that the order of operations is Parenthesis, Exponents, Multiplication and Division, Addition and Subtraction (PEMDAS). A review of order of operations can be found in Appendix A.

WORKED EXERCISES

Temperature Conversions

1-20 While traveling in Switzerland, you feel ill and visit a health clinic. The nurse informs you that you have a temperature of $38.2\ ^\circ\mathrm{C}$. Convert this temperature into degrees Fahrenheit. Do you have a fever?

Solution:

First, we determine that we have been *given* a temperature in $^\circ\mathrm{C}$ and *asked for* the temperature in $^\circ\mathrm{F}$. Thus, we need a $\mathrm{T}_{^\circ\mathrm{C}}$ to $\mathrm{T}_{^\circ\mathrm{F}}$ conversion. Next, we determine that the temperature equation we need from Table 1-5 is the equation in the first column:

$$T_{^\circ\mathrm{C}} = \frac{(T_{^\circ\mathrm{F}} - 32.0)}{1.8}$$

Next, we determine if the equation is expressed in terms of the *asked for* unit. If not, we need to rearrange it algebraically, so that it is. Since the equation does not read $\mathrm{T}_{^\circ\mathrm{F}} = \ldots$, we need to rearrange the equation algebraically so that $\mathrm{T}_{^\circ\mathrm{F}}$ is isolated on one side of the equal sign.

$$T_{^\circ\mathrm{F}} = (1.8 \times T_{^\circ\mathrm{C}}) + 32.0$$

Next, we substitute the given temperature, $38.2\ ^\circ\mathrm{C}$, for $\mathrm{T}_{^\circ\mathrm{C}}$ and solve the equation:

$$\begin{aligned} T_{^\circ\mathrm{F}} &= (1.8 \times 38.2) + 32.0 \\ &= 100.8\,^\circ\mathrm{F} \end{aligned}$$

Normal body temperature is $98.6\ ^\circ\mathrm{F}$ so you do have a fever. Significant figure rules for addition and subtraction apply. Since the *given* temperature has one place to the right of the decimal, the final answer should be rounded to the tenths place. (The 1.8 in the equation is treated as an exact number.)

1-21 Helium is a gas at room temperature, making it convenient for filling balloons. Helium is a liquid at $4\ \mathrm{K}$. Convert this temperature into $^\circ\mathrm{F}$.

Solution:

First, we determine that the *given* temperature is in kelvins $(4\ \mathrm{K})$ and that we are being *asked for* the temperature in $^\circ\mathrm{F}$. We have not been given an equation that converts between K and $^\circ\mathrm{F}$, therefore, we have to perform two consecutive temperature conversions, first from T_K to $\mathrm{T}_{^\circ\mathrm{C}}$ and then one from $\mathrm{T}_{^\circ\mathrm{C}}$ to $\mathrm{T}_{^\circ\mathrm{F}}$. Therefore, we need both equations in Table 1-5 or algebraically rearranged forms of them.

$$T_{^\circ C} = T_K - 273.15 \quad \text{and} \quad T_{^\circ\mathrm{C}} = \frac{(T_{^\circ\mathrm{F}} - 32)}{1.8}$$

$$\text{or} \quad T_{^\circ\mathrm{F}} = (1.8 \times T_{^\circ\mathrm{C}}) + 32.0$$

Step 2: The first temperature conversion requires conversion of the *given* temperature, in K, into $^\circ\mathrm{C}$, which requires the equation above left. Since the equation is already expressed in terms of $\mathrm{T}_{^\circ\mathrm{C}}$, no algebra is required and we can use the equation as is. We substitute the *given* temperature into T_K:

$$\begin{aligned} T_{^\circ\mathrm{C}} &= T_K - 273.15 \\ &= 4 - 273.15 \\ &= -269^\circ\mathrm{C} \end{aligned}$$

Step 3: In the second temperature conversion, we need to convert the value from step 2 into the *asked for* units of $^\circ\mathrm{F}$, which requires the equation in step 1. Since the equation is expressed in terms of $\mathrm{T}_{^\circ\mathrm{C}}$, we use the algebraically rearranged equation solved for $\mathrm{T}_{^\circ\mathrm{C}}$. We then substitute the temperature calculated in step 2, $-269^\circ\ \mathrm{C}$, into the equation and solve:

$$T_{^\circ F} = (1.8 \times T_{^\circ C}) + 32.0$$

$$T_{^\circ F} = (1.8 \times -269) + 32.0$$

$$= -452\ ^\circ F$$

PRACTICE EXERCISES

35. Many birds have a normal body temperature of $106\ ^\circ F$. Convert this temperature into $^\circ C$ and kelvins.

36. A disoriented patient comes into the emergency room after running the Los Angeles Marathon complaining of nausea and cramps. You check her temperature and find that she has a core temperature of $39\ ^\circ C$. Convert this temperature to degrees Fahrenheit. Could she be suffering from a heat stroke?

Chemistry in Medicine

Matter, Energy, and Starvation

According to the Food and Agriculture Organization (FAO) of the United Nations, 795 million people (11% of the world's population) were undernourished in 2015, down by 167 million over the past decade. Food is matter, and as we learned in this chapter, energy and matter are closely linked. Energy is defined as the ability to do work. Your body does work as it performs basic involuntary physiological functions, such as breathing, keeping the heart beating, repairing damaged cells, and so forth. Your body obtains energy from the food you eat, which allows this cellular work to occur, as well as activities such as walking, talking, and studying. The carbohydrates and fats in our food are the main source of energy for our cells, which require a constant supply of energy in order to perform their work. Proteins can provide energy but they are primarily used to build cell components.

Carbohydrates and fats are a type of matter that is high in potential energy. As part of the digestive process, carbohydrates are broken down into the sugar glucose, which then enters the bloodstream (blood sugar) and is delivered to cells throughout the body. The body also breaks down glycogen, its storage form of glucose, when energy is needed. Our glycogen stores typically last for 24 hours, depending on our activity level, and sustain us between meals. Glucose is the preferred energy source for most cells, and the only source of energy for brain cells—neurons.

When glucose enters a cell, several chemical changes occur. Chemical changes—chemical reactions—that occur in living systems are referred to as metabolism. The metabolism of glucose requires oxygen, which is why we need to breathe. During the metabolism of glucose, potential energy from glucose is transferred to the cell, which uses the energy to perform work. Proteins can be converted into glucose, but it *costs* energy to convert proteins into glucose—in fact, it costs as much energy as the glucose supplies in energy. Thus, the body converts proteins into glucose only when it is starving. Per gram, fats provide almost twice as much energy as carbohydrates; however, fats cannot be converted into glucose—the source of energy required by our brain cells.

Starvation occurs when the body is not consuming enough food to produce the energy it needs to perform the basic cellular *work* of living—the functions needed to stay alive (**Figure 1-20**). The body then turns to its own matter for energy. We are familiar with the loss of fat during starvation, as the body turns to its stored energy—fat—to perform work. Since fat is not a source of glucose, however, the body will also metabolize the protein in muscle to supply energy for the brain, including the heart muscle. In the process, the body withers away trying to keep itself alive. Starvation for one or more months ultimately results in death. Even short-term periods of starvation can have severe health consequences. While fewer than 2.5% of Americans are undernourished, health care professionals encounter undernourishment in patients with eating disorders, mental illness, the elderly, the poor, and even some college students. Starvation is still a major problem in developing countries—especially for children (Figure 1-20).

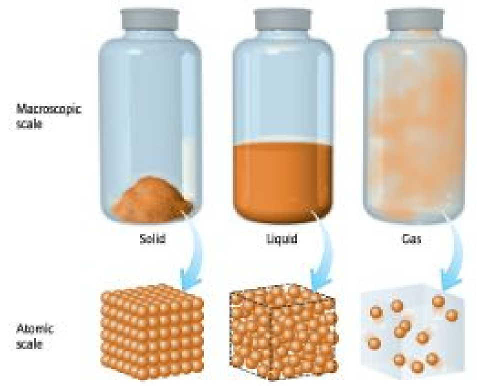
The Times/Gallo Images/Getty Images

Figure 1-20 Every year 3.1 million children around the world die from starvation.

Core Concepts

1.1 Matter and Energy

- In the solid state, particles of matter are in an ordered arrangement and interacting. In the liquid state, particles are disordered but interacting with other particles. In the gas state, particles are far apart and not interacting.
- A faster moving object has more kinetic energy than a slower moving object.
- For a given substance, particles in the gas state have more kinetic energy than particles in the liquid state, which have more kinetic energy than particles in the solid state.
- Temperature is a measure of the average kinetic energy of particles of matter. For a given substance, the gas state has a higher temperature than the liquid state, and the liquid state has a higher temperature than the solid state.
- Heat is the kinetic energy transferred from matter at a higher temperature to matter at a lower temperature.
- Matter has potential energy as a result of position or composition. Some substances are higher in potential energy (less stable) while other substances are lower in potential energy (more stable).



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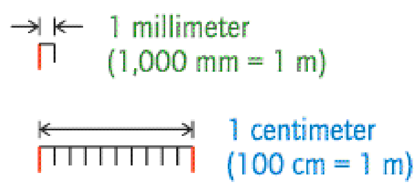
1.2 Units and Measurement

- A measurement consists of a numerical value and a unit.
- The metric system is built on a set of base units, each denoted by a one-letter symbol.
- A prefix attached to a base unit increases or decreases the numerical value by a factor of 10^{x} or 10^{-x}.



- To create a conversion between a prefixed unit and its base unit, set the prefixed unit equal to the multiplier times the base unit: $\text{prefixed unit} = (10^{x \text{ or } -x}) \times (\text{base unit})$
- A metric prefix can be used with any base unit, and always represents the same multiplier.
- Volume is a unit of measurement derived from length where $1 \text{ mL} = 1 \text{ cm}^3$ and $1 \text{ L} = 1{,}000 \text{ cm}^3$.
- The volume of an object can be measured by determining the volume of water that it displaces: $V_{object} = V_{final} - V_{initial}$

1.3 Significant Figures in Measurements and Calculations

- Uncertainty in a measurement is conveyed through the number of significant figures reported: The last significant figure in a measurement is the uncertain and estimated digit.
- Zeros that serve as placeholders in a number are not significant figures.
- Exact numbers are not measurements. They include numbers obtained by accurate counting and defined conversions.
- In calculations involving multiplication or division of measured values, round the final answer such that it has the same number of significant figures as the measurement with the fewest number of significant figures.

(a) (b)

- In a calculation involving addition or subtraction of measured values, the answer cannot have more places to the right of the decimal than the measurement with the fewest places to the right of the decimal.

1.4 Dimensional Analysis

- Unit conversions—converting a measurement from one unit into another unit—are reliably solved using dimensional analysis.

$$5.0\ \text{cm}^3 \times \frac{19.32\ \text{g}}{1\ \text{cm}^3} = 97\ \text{g}$$

- Dimensional analysis allows us to multiply two or more conversion factors, setting them up in sequence such that units cancel, until only the asked for unit remains.

- When the given unit and asked for unit are from different systems of measurement, an Englishmetric conversion is identified and used as one of the conversions using dimensional analysis.

- Density is defined as the mass divided by the volume of a sample of material. Density can be used as a conversion factor to calculate the mass or volume of a substance when given the density and either the volume or mass of the sample.

- Specific gravity is a unitless measurement defined as the density of a substance divided by the density of water $(1.00\ \text{g/mL})$.

- When dosage is given as a mass of medication per mass of body weight, it can be expressed as a ratio, which can be used as a conversion factor in dosage calculations using dimensional analysis.

1.5 Temperature Scales and Conversions

- Three temperature scales—Fahrenheit $(^\circ\text{F})$, Celsius $(^\circ\text{C})$, and Kelvin (K)—are used to report temperature. The Kelvin scale is an absolute temperature scale.

- Two equations are used to interconvert between the temperatures scales. One equation is used for interconverting between $^\circ\text{C}$ and $^\circ\text{F}$ and another equation is used for interconverting between $^\circ\text{C}$ and K.

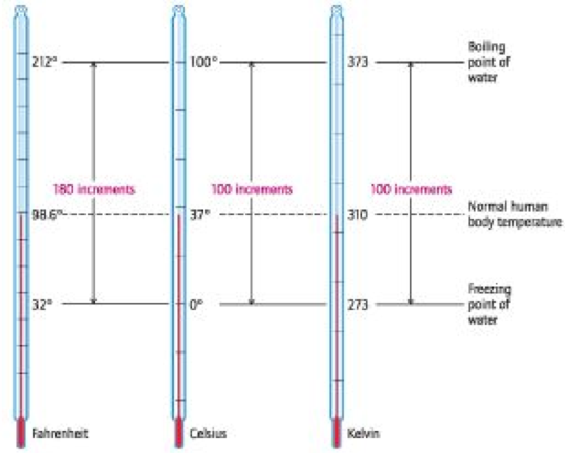
Key Words

Atomic scale:
The scale of atoms and molecules, a size so small that we cannot see it even with a light microscope.

Base unit:
The basic units of measurement in the metric system, such as the gram (g) for mass, the meter (m) for length, and the liter (L) for volume.

b.i.d.:

Latin abbreviation indicating a medication should be administered twice daily



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Celsius scale:

The temperature scale used in science, medicine, and for everyday applications in most of the world outside the United States. The freezing point of water is defined as $0\ ^\circ\mathrm{C}$ and the boiling point of water as $100\ ^\circ\mathrm{C}.$

Conversion:

A mathematical expression that equates two different units.

Conversion factor:

Expression of a conversion as a ratio, wherein one side of the equality appears in the numerator and the other side of the equality appears in the denominator. There are two possible representations.

Density:

A physical property of a substance defined as its mass (*m*) divided by its volume $(V):\ d = m/V.$

Dimensional analysis:

A process for solving problems that focuses on the units to direct how a calculation is set up. One or more conversion factors are used so that units cancel until all but the asked for unit remains.

Energy:

The capacity to do work, where work is the act of moving an object against an opposing force.

English system:

A system of measurement used in the United States, primarily for nonscientific applications. Includes units like the pound (lb), the gallon (gal), and the foot (ft).

Exact numbers:

Numbers with no limit in their number of significant figures because they are not measured values. They include numbers obtained by accurate counting, or defined quantities.

Fahrenheit scale $(^\circ\mathbf{F})$:

A temperature scale used in the United States for everyday applications. The freezing point of water is defined as $32\ ^\circ\mathrm{F}$ and the boiling point of water as $212\ ^\circ\mathrm{F}.$

Gas:

A state of matter where particles of matter are very far apart, interacting only during the occasional collision. For a given substance, the gas state has the greatest kinetic energy.

Gram $(\mathbf{g})$:

The metric base unit of mass and weight.

Heat:

The kinetic energy transferred from a sample of matter at a higher temperature to a sample of matter at a lower temperature. Heat always flows in the direction hot to cold.

Hypothermia:

When the core temperature of the body falls below $35\ ^\circ\mathrm{C}\ \ (95\ ^\circ\mathrm{F}).$

Kelvin scale:

The absolute temperature scale used in the sciences. $0\ \mathrm{K}$ represents the temperature at which particles have no kinetic energy.

Kinetic energy:

The energy of motion; the energy a substance has as a result of the motion of its particles.

Length:

The distance between two points. The base unit of length in the metric system is the meter. The SI unit of length is the meter. In the English system, the common units of length are the mile, the yard, the foot, and the inch.

Liquid:

A state of matter where particles of matter are disordered but still interacting with each other. For a given substance, the liquid state has more kinetic energy than the solid state but less than the gas state.

Liter (L):
The metric base unit of volume.

Mass:
A measure of the amount of matter. The base unit of mass in the metric system is the gram. The SI unit of mass is the kilogram, kg. The common English units of mass are the pound (lb) and the ounce (oz).

Macroscopic scale:
Matter that can be seen with the naked eye.

Matter:
Anything that has mass and occupies space (volume).

mcg:
An alternative abbreviation for the microgram, used in nutritional applications. $1\ \mathrm{mcg} = 10^{-6}\ \mathrm{g}.$

Meniscus:
The curved surface formed by a liquid in contact with glass. Reading a volume measurement should be taken from the *bottom* of the meniscus.

Meter (m):
The metric base unit of length and distance.

Metric prefix:
The abbreviation preceding a metric base unit that acts as a multiplier to increase the numerical value by a factor of 10^{x} or decrease the numerical value by a factor of $10^{-\mathrm{x}}$. For example, *milli* (m) is the prefix with the multiplier 10^{-3}.

Metric system:
A system of measurement used in science, medicine and in most of the world. It is based on a set of base units that can be combined with various prefixes to increase or decrease the base unit.

Microscopic scale:
Matter too small to be seen with the naked eye but that can be seen under a light microscope.

Potential energy:
Stored energy. The energy a substance has as a result of its position or composition.

Prefix:
See Metric prefix.

q.d.:
Latin abbreviation indicating a medication should be administered once daily.

Second (s):
The metric base unit of time.

SI System:
An International System of Units that is based on the metric system and establishes a uniform set of units in science and commerce.

Significant figures:
All the certain digits plus one estimated digit used when recording a measurement. The number of significant figures conveys the degree of uncertainty in a measurement.

Solid:
A state of matter where particles of matter are close together and in a very ordered arrangement. For a given substance, the solid state has the least amount of kinetic energy.

Specific gravity:
A measurement related to density used for special applications. Specific gravity, a unitless quantity, is the density of a substance divided by the density of water $(1.0\ \mathrm{g/mL})$.

States of matter:
The physical forms in which matter is found: solid, liquid, and gas.

Substance:
A general reference to a specific type of matter.

Temperature:
A measure of the average kinetic energy of the particles of a sample of matter. The three temperature scales, Celsius, Fahrenheit, and Kelvin, are used to measure temperature.

Unit:
The part of a measurement that indicates the type of quantity measured. Units are part of the metric or the English system.

Unit conversion:

A type of calculation in which a measurement in one unit is converted into the equivalent value in another unit. The unit can be from the same or different systems of measurement.

Volume:

A measure of three-dimensional space occupied by a substance. Volume is a unit of measure derived from units of length. The liter is the base unit of volume in the metric system. Both the cm^3 and the mL are equivalent metric units of volume. Common English units of volume include the gallon, the pint, and the quart.

Work:

The act of moving an object against an opposing force.

Additional Exercises

Osteoporosis and Measurement of Bone Density

37. How is bone strength assessed?

38. Does a strong bone have more or less bone mineral per volume than a weak bone?

39. What type of scan is used to estimate BMD?

40. How is a T-score assigned?

41. A 50-year-old woman had a DEXA scan taken. Her T-score is $-1.3.$ What is the condition of her bone: normal, osteopenia, or osteoporosis?

42. A 65-year-old woman had a DEXA scan taken, and her T-score is $+0.9.$ What is the condition of her bone: normal, osteopenia, or osteoporosis?

1.1 Matter and Energy

43. What are the three physical states of matter and their abbreviations?

44. In which two states of matter do particles interact with other particles?

45. In which two states of matter are the particles the farthest apart?

46. Is heat, temperature, or work a measure of the average kinetic energy of particles?

47. Is heat, temperature, or work involved in moving an object?

48. Does a ball sitting at the top of a hill have potential energy or kinetic energy?

49. Does a ball rolling down a hill have potential energy or kinetic energy?

50. Indicate whether each of the following examples is a demonstration of potential energy or kinetic energy:
    a. water flowing over a dam
    b. a skier standing at the top of a hill
    c. a dancer dancing
    d. the sandwich you ate for lunch

51. Indicate whether each of the following examples is a demonstration of potential energy or kinetic energy:
    a. a biker pedaling up a hill
    b. a hiker standing at the top of a mountain
    c. the helium atoms in a balloon
    d. the wax in a candle

52. In which physical state do molecules have the least amount of kinetic energy?
    a. solid state
    b. liquid state
    c. gas state

53. In which physical state do molecules have the greatest amount of kinetic energy?
    a. solid state
    b. liquid state
    c. gas state

54. In which state of matter do water molecules have the most kinetic energy: liquid water, steam, or ice?

1.2 Units and Measurement

55. Classify the size of the following examples of matter as being on the macroscopic, microscopic, or the atomic scale:

a. a hospital
b. a skin cell
c. DNA
d. a red blood cell

56. Classify the size of the following examples of matter as being on the macroscopic, microscopic, or the atomic scale:
a. a lead atom
b. the human body
c. a grain of sand
d. a virus

57. What is the SI system?

58. What is the SI unit for mass?

59. What quantity does the gram measure?

60. What quantity does the meter measure?

61. What is the metric base unit for volume?

62. What is the metric base unit for time?

63. Indicate the multiplier that each of the metric prefixes represents, and then order them from smallest metric unit to largest metric unit when placed before a base unit (for example $\text{centi} < \text{deci}$):
a. nano
b. kilo
c. pico
d. micro

64. Indicate the multiplier that each of the metric prefixes represents, and then order them from smallest metric unit to largest metric unit when placed before a base unit (for example $\text{centi} < \text{deci}$):
a. giga
b. centi
c. milli
d. deci

65. Which of the following measurements have English units and which have metric units? Also, identify each as a unit of length, mass, or volume:
a. $25.4\ \text{mm}$
b. $2.54\ \text{in.}$
c. $33.25\ \text{oz}$
d. $454.1\ \text{g}$

66. Which of the following measurements have English units and which have metric units? Also, identify each as a unit of length, mass, or volume.
a. $2.12\ \text{kg}$
b. $2.54\ \text{lb}$
c. $333\ \text{mL}$
d. $12.5\ \text{ft}$

67. Write the conversion between centimeters and meters, using the symbols for these units.

68. Write the conversion between picoseconds and seconds, using the symbols for these units.

69. Write the conversion between megabytes and bytes, using the symbols for these units. The symbol for a byte is B.

70. Write the conversion between a kilowatt and a watt. The symbol for a watt is W.

71. Which is larger: $4.5 \times 10^{-2}\ \mu\text{g}$ or $4.5 \times 10^{2}\ \mu\text{g}$? Is the microgram a measure of length, mass, or volume?

72. Which is smaller: $6.3 \times 10^{-3}\ \text{mm}$ or $6.3 \times 10^{3}\ \text{mm}$? Is the millimeter a measure of length, mass, or volume?

73. Name three types of specially marked glassware that are used to measure liquids.

74. What is a meniscus and when is it formed?

75. When you are reading the volume from a graduated cylinder, where would the volume be read?
   a. At the highest point of the liquid
   b. At the middle of the meniscus
   c. At the bottom of the meniscus

76. A lead ball is added to a graduated cylinder containing $15.0\ \mathrm{mL}$ of water, causing the level of the water to increase to $16.5\ \mathrm{mL}.$ What is the volume in milliliters of the lead ball?

77. An irregular-shaped metal object is placed in a graduated cylinder containing $200.\ \mathrm{mL}$ of water. The water level increases to $203.5\ \mathrm{mL}.$ What is the volume in milliliters of the metal object?

78. What is the volume, in milliliters, of a cube measuring $24\ \mathrm{cm}$ per side?

79. What is the volume, in $\mathrm{cm}^3,$ of a cube measuring $5.21\ \mathrm{cm}$ per side?

1.3 Significant Figures in Measurements and Calculations

80. For each of the following, indicate whether it is an exact number or a measured value. If it is a measured value, indicate the number of significant figures and underline the estimated digit:
   a. $57{,}000\ \mathrm{m}$
   b. $4.60\ \mathrm{mL}$
   c. $0.00011\ \mathrm{g}$
   d. $23{,}304.60\ \mathrm{s}$
   e. $256\ \mathrm{nurses}$

81. For each of the following, indicate whether it is an exact number or a measured value. If it is a measured value, indicate the number of significant figures and underline the estimated digit:
   a. $304\ \mathrm{mm}$
   b. $429\ \mathrm{bees}$
   c. $5{,}110\ \mathrm{minutes}$
   d. $0.000330\ \mathrm{kg}$
   e. $5{,}000.\ \mathrm{g}$

82. Round the following measurements to 3 significant figures:
   a. $2.30653\ \mu\mathrm{g}$
   b. $9{,}3129\ \mathrm{mm}$
   c. $1.555\ \mathrm{L}$
   d. $5678.9\ \mathrm{seconds}$

83. Round the following measurements to 2 significant figures:
   a. $1.7777\ \mathrm{nm}$
   b. $4.25\ \mathrm{mL}$
   c. $28.1\ \mathrm{pg}$
   d. $357.8\ \mathrm{m}$

84. Perform the following calculation: $\frac{3.27}{5.2}.$ Assume these are measured values. Which is the correct answer?
   a. 0.6288

b. 0.629
c. 0.63
d. 0.6

85. Perform the following calculation: $124.893–45.01.$ Assume these are measured values. Which is the correct answer?
a. 79.9
b. 79.88
c. 79.883
d. 80

86. Perform the following calculations and report the correct number of significant figures. Assume these are measured values.
a. $3.2 \times 8.54 =$
b. $3.2 + 8.54 =$

87. Perform the following calculations and report the correct number of significant figures. Assume these are measured values.
a. $2.26 + 8.1 =$
b. $2.26 \times 8.1 =$

88. Perform the following calculations. Show the correct number of significant figures in your answer, assuming each number is a measured value. Include units in your answers.
a. $56.33 \text{ cm} \times 2.50 \text{ cm} =$
b. $3.4 \text{ cm} + 2.2 \text{ cm} + 5.11 \text{ cm} + 8.777 \text{ cm} =$
c. $\frac{33.22 \text{ g}}{39.0 \text{ mL}} =$

89. Perform the following calculations. Show the correct number of significant figures in your answer, assuming each number is a measured value.
a. $33{,}000. + 910. =$
b. $0.333 \times 0.22 =$
c. $(37.55 \text{ mL} + 22.2 \text{ mL}) \times 5.666 =$

1.4 Dimensional Analysis

90. Perform the following metric conversions and report your final answer to the correct number of significant figures.
a. Convert $50{,}000 \text{ meters}$ into kilometers.
b. Convert 0.66 grams into micrograms.

91. Perform the following metric conversions and report your final answer to the correct number of significant figures.
a. How many milliliters is equivalent to $6.0 \text{ L}?$
b. How many kilometers is equivalent to $2.0 \times 10^6 \text{ m}?$

92. Which of the following is equivalent to $1 \text{ m}?$
a. 10 cm
b. 100 dm
c. $1{,}000 \text{ mm}$
d. $10^{12} \ \mu\text{m}$

93. Which of the following is equivalent to $1 \text{ g}?$
a. 10 kg
b. 0.001 kg

c. $100\ \text{mg}$
d. $10^8\ \mu\text{g}$

94. What conversion factor would you use to convert $4{,}000\ \text{m}$ into kilometers?

95. A doctor must make an incision $2.5\ \text{cm}$ long. What is the length of this incision in meters?

96. Ibuprofen can be found in $200\ \text{mg}$ doses in over-the-counter analgesics such as Advil and Motrin. How many grams of ibuprofen does such a tablet contain?

97. Convert $500.\ \text{mg}$ of vitamin C into grams (g) and micrograms (μg).

98. What two conversion factors would you use to convert $61{,}000\ \text{mm}$ into picometers?

99. How many milliliters is equivalent to $75.6\ \mu\text{L}$?

100. What is the weight, in pounds, of an animal that weighs $150\ \text{kg}$?

101. How many kilometers is equivalent to $68.2\ \text{miles}$?

102. How many liters is $86\ \text{gallons}$?

103. An ultrasound technician measures the humerus bone in an 18-week-old fetus. The bone measures $26.7\ \text{mm}$ in length. What is the length of the bone in inches?
a. $6.78 \times 10^{-3}\ \text{in.}$
b. $1.05\ \text{in.}$
c. $1.05 \times 10^{3}\ \text{in.}$
d. $6.78\ \text{in.}$

104. A phlebotomist draws two tubes of blood for a total volume of $14.3\ \text{cm}^3$. How many liters of blood are in the two tubes?
a. $14.3\ \text{L}$
b. $1.43\ \text{L}$
c. $1.43 \times 10^{3}\ \text{L}$
d. $1.43 \times 10^{-2}\ \text{L}$

105. Convert the following lengths into millimeters:
a. $1\ \text{km}$
b. $1\ \text{cm}$
c. $1\ \text{dm}$
d. $1\ \text{nm}$

106. Convert the following masses into kilograms:
a. $1\ \text{g}$
b. $1\ \text{ng}$
c. $100\ \text{mcg}$
d. $10\ \mu\text{g}$

107. Convert the following volumes into microliters:

a. $5\ \mathrm{mL}$
b. $0.5\ \mathrm{L}$
c. $250\ \mathrm{cm}^3$
d. $80\ \mathrm{dL}$

108. Which of the following is equivalent to $150\ \mu\mathrm{g}$?

a. $1.5 \times 10^{-5}\ \mathrm{g}$
b. $1.5 \times 10^{-2}\ \mathrm{mg}$
c. $1.5 \times 10^{-7}\ \mathrm{kg}$
d. $150\ \mathrm{g}$

109. A sample of muscle tissue has a volume of $8.7\ \mathrm{mL}$ and a mass of $9.22\ \mathrm{g}$. What is the density of the muscle tissue?

110. A sample of compact bone has a mass of $3.8\ \mathrm{g}$ and a volume of $2.0\ \mathrm{cm}^3$. What is the density of the sample?

111. What substance has a density of $1.0\ \mathrm{g/mL}$?

112. If a liquid that does not mix with water has a density greater than $1.0\ \mathrm{g/mL}$, will it float or sink in water?

113. Using Table 1-4, calculate the mass of a gold sphere that displaces $2.3\ \mathrm{mL}$ of water?

114. Using Table 1-4, calculate the mass of a gold cube having sides $2.20\ \mathrm{cm}$ in length.

115. The density of a patient's urine is $1.025\ \mathrm{g/mL}$. What is the specific gravity of the urine sample? Is the specific gravity within the normal range? The normal specific gravity of urine is $1.002–1.030$.

116. A patient provides a urine sample. The density of the patient's urine is $1.037\ \mathrm{g/mL}$. What is the specific gravity of the urine sample? Is the specific gravity within the normal range? The normal specific gravity of urine is $1.002–1.030$.

117. A patient provides a urine sample. The specific gravity of the patient's urine is 1.014. What is the density of the urine sample? Is the specific gravity within the normal range? The normal specific gravity of urine is $1.002–1.030$.

118. The specific gravity of a patient's urine is 0.997. What is the density of the patient's urine? Is the specific gravity within the normal range? The normal specific gravity of urine is $1.002–1.030$.

119. The initial dose of Accupril, used to treat hypertension, is $10\ \mathrm{mg}$ *q.d.* How many times a day should Accupril be administered? How many milligrams should be given at every administration?

120. Retrovir is one of the drugs used to treat HIV. The recommended dose is $600\ \mathrm{mg}$ per day *b.i.d.* How many times a day should Retrovir be administered? How many milligrams should be given at every administration?

121. Tylenol is ordered for a child with a fever at a dose of $25\ \mathrm{mg}$ per kg of body weight a day. If the child weighs $34\ \mathrm{lb}$, what mass of Tylenol, in milligrams, should be given to the child every day?

122. Benadryl is used to treat itchy skin in dogs. The recommended dose is 1 milligram per pound. What mass of Benadryl, in milligrams, should be given to a dog that weighs $27\ \mathrm{kg}$?

123. The recommended dose of Ceclor, an antibiotic, is 20 mg per kg of body weight a day, in three equally divided doses (every 8 hours). How many milligrams should a baby weighing 12 lb receive at each administration?

124. A 16 lb infant has been exposed to the flu virus in the community. The recommended dose of a prophylactic Tamiflu treatment is 3 mg per kg *b.i.d.* How many times a day should Tamiflu be administered? How many milligrams of Tamiflu should be given at every administration?

1.5 Temperature Scales and Conversions

125. What is the freezing point of water in degrees Celsius and degrees Fahrenheit?

126. What is normal body temperature in degrees Celsius and degrees Fahrenheit?

127. Benzamycin is a topical antibiotic gel. The gel needs to be stored between 2 °C and $8 \text{ °C}.$ Which of the following would be the best place to store this gel?
   a. in the freezer
   b. in the refrigerator
   c. in a medicine cabinet
   d. outdoors

128. Accutane is used to treat acne. Accutane gelatin capsules need to be stored between 15 °C and $30 \text{ °C}.$ Where should these capsules be stored?
   a. in the freezer
   b. in the refrigerator
   c. in a medicine cabinet
   d. outdoors

129. While traveling in Europe, you notice an outdoor thermometer that reads $18 \text{ °C}.$ Are you more likely to be wearing winter clothes (coat, hat, scarf, etc.) or summer clothes? What is this temperature in kelvins and in degrees Fahrenheit?

130. A thermometer in Christchurch, New Zealand, reads $31\text{°C}.$ Are you wearing winter clothes (coat, hat, scarf, etc.) or summer clothes? What is this temperature in degrees Fahrenheit and kelvins?

131. A patient is experiencing heat stroke. Her body temperature is $105.2 \text{ °F}.$ What is this temperature in degrees Celsius?

132. A patient has moderate hypothermia. His body temperature is $87.4 \text{ °F}.$ What is this temperature in degrees Celsius?

133. A patient has a body temperature of $33 \text{ °C}.$ Does this patient have heat stroke or hypothermia? What is this temperature in degrees Fahrenheit?

134. The temperature in outer space is $2.7 \text{ K}.$ What is this temperature in degrees Celsius and in degrees Fahrenheit?

135. Medical couriers use dry ice to keep perishable medical materials cold while they are transported. Dry ice has a temperature of $-78 \text{ °C}.$ What is this temperature in kelvins and degrees Fahrenheit?

136. Liquid nitrogen is used to freeze off warts. Liquid nitrogen has a temperature of $77 \text{ K}.$ What is this temperature in degrees Celsius and degrees Fahrenheit?

Chemistry in Medicine: Matter and Energy in Malnutrition

137. From what type of foods do we obtain most of our energy: carbohydrates, proteins, or fats?

138. What substance, circulating in our blood, is the most important source of energy for our cells?

139. When the body is starving, how does it produce energy for brain cells?

140. Is glucose (blood sugar) high or low in potential energy?

141. Is carbon dioxide high or low in potential energy?

Challenge Questions

142. What is the length of the sides of a cube with a volume of $8 \text{ cm}^3?$

143. Rank the following volume measurements from largest to smallest: $50.00\ \text{mL};\ 5{,}000\ \mu\text{L};\ 0.5000\ \text{L};\ 8.000\ \text{cm}^3$.

144. For each of the following pairs, indicate which length is shorter? If they are the same, state so.
   a. $10\ \text{m or } 1\ \text{km}$
   b. $1\ \text{nm or } 10^{-9}\ \text{m}$
   c. $10^{-3}\ \text{m or } 1\ \text{mm}$
   d. $1\ \text{nm or } 1\ \mu\text{m}$

145. For each of the following pairs, indicate which measurement represents the smaller mass. If they have the same mass, state so.
   a. $1\ \text{ng or } 1\ \text{mg}$
   b. $100\ \text{mg or } 1\ \text{g}$
   c. $1{,}000\ \text{mg or } 1\ \text{g}$
   d. $50\ \text{mcg or } 100\ \mu\text{g}$

146. What is the mass, in milligrams, of a cube of iron that has sides measuring $2.5\ \text{inches}$ in length?

147. A $62\ \text{lb}$ child has a *streptococcus* infection. Amoxicillin is prescribed at a dosage of $25\ \text{mg}$ per kg of body weight *b.i.d.* Amoxicillin should be stored at or below 20°C.
   a. What is the meaning of the Latin abbreviation *b.i.d.*?
   b. How often should amoxicillin be administered to the child?
   c. How many milligrams of amoxicillin should be given at each administration?
   d. Should the amoxicillin be stored in the freezer or the refrigerator?
   e. Amoxicillin is available as a tablet or powder. Are the particles in the tablet or powder close together or far apart?

148. What is the volume, in milliliters, of a cube measuring $3.2\ \text{mm}$ per side?

149. If a cube has a volume of $147\ \text{cm}^3$, what are the lengths of the sides of this cube in inches?

Answers to Practice Exercises

1. Liquid water flows when poured, steam does not. Liquid water does not occupy the entire volume of its container, whereas steam does. Steam is hotter than liquid water. The water particles (molecules) are close together and interacting with one another in liquid water, whereas those same particles are far apart and only occasionally interacting in the gas phase. Steam has more kinetic energy than liquid water, hence it has a higher temperature.

2. a. potential energy
   b. kinetic energy
   b. kinetic energy
   d. kinetic energy
   e. potential energy

3. Heat is a type of kinetic energy.

4. $1\ \text{mL} = 10^{-3}\ \text{L}.\ \ 1\ \text{dL} = 10^{-1}\ \text{L}$

5. $1\ \mu\text{g} = 10^{-6}\ \text{g}.\quad 1\ \mu\text{m} = 10^{-6}\ \text{m}.\quad 1\ \mu\text{s} = 10^{-6}\ \text{s}$

6. a. $1\ \text{mm} = 10^{-3}\ \text{m}$
   b. $1\ \text{dm} = 10^{-1}\ \text{m}$
   c. $1\ \text{km} = 10^{3}\ \text{m}$

7. $1 \text{ mi} = 5{,}280 \text{ ft};$ these are units of distance in the English system.

8. a. μm is a metric unit of length
   b. gal is an English unit of volume
   c. kg is a metric unit of mass
   d. cm^3 is a metric unit of volume
   e. lb is an English unit of mass
   f. mL is a metric unit of volume

9. a, b, and d would be visible to the naked eye.

10. $11 \text{ mL} – 10. \text{ mL} = 1 \text{ mL} = 1 \text{ cm}^3$

11. $2 \text{ cm} \times 2 \text{ cm} \times 2 \text{ cm} = 8 \text{ cm}^3 = 8 \text{ mL}$

12. a. Measured value; 1 significant figure $0.00\underline{7}$
    b. Exact number, obtained by counting the people
    c. Measured value; 5 significant figures because there is a decimal after the zeros: $2\underline{3},000.$
    d. Measured value; 4 significant figures $0.00405\underline{0}$
    e. Measured value; 2 significant figures $3\underline{2}00,$ the zeros are placeholders because there is no decimal after the last zero.

13. a. i. $5.\underline{5}$
    ii. $5.5\underline{2}$
    iii. $5.509\underline{3}$
    b. ii.
    c. i. 2
    ii. 3
    iii. 5

14. a. $2{,}146 \text{ m}^2$
    b. 15.21 g
    c. 3.2 mL

15. $77 \text{ kg};$ kg is more convenient than g because the numerical value has fewer zeros.

16. 0.561 L

17. $5.5 \times 10^{-9} \text{ m}$

18. 0.22 g

19.

$$150.\ \cancel{\mu\text{m}} \times \frac{10^{-6}\ \cancel{\text{m}}}{1\ \cancel{\mu\text{m}}} \times \frac{1 \text{ mm}}{10^{-3}\cancel{\text{m}}} = 0.150 \text{ mm}$$

20.

$$25\ \cancel{\text{dL}} \times \frac{10^{-1}\ \cancel{\text{L}}}{1\ \cancel{\text{dL}}} \times \frac{1 \text{ mL}}{10^{-3}\cancel{\text{L}}} = 2500 \text{ mL}$$

21.

$$1.1\ \cancel{\text{km}} \times \frac{10^3\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1\ \text{cm}}{10^{-2}\cancel{\text{m}}} = 110{,}000\ \text{cm};$$

therefore $520{,}000\ \text{cm}$ is the longer distance.

22. $57\ \text{L}$

23. $2.00\ \text{lb}$

24.

$$285\ \text{pts} \times \frac{1\ \text{qt}}{2\ \text{pt}} \times \frac{1\ \text{L}}{1.06\ \text{qt}} = 134\ \text{L}$$

25. $1.1\ \text{g/mL}$

26. $d = m/V;\ d = 71.1\ \text{g}/9.0\ \text{cm}^3 = 7.9\ \text{g/cm}^3,$ which is the density of iron so the unknown metal is iron.

27. $d = m/V;\ d = 10.5\ \text{g}/\left(95.0\ \text{mL} - 82.5\ \text{mL}\right) = 0.840\ \text{g/mL}$

28. $1.1\ \text{cm}^3$

29. $25\ \text{g}$

30. A woman with osteoporosis has lost bone mass; therefore, her bone density (m/V) would be less than normal.

31. Since ice floats on water, it has a density less than liquid water.

32.
   a. $\frac{25.0\ \text{mg}}{1\ \text{kg}}$
   b. twice daily; every $12\ \text{hours}$
   c. $624\ \text{mg}$ twice a day

33.
   a. The medication should be given every $6\ \text{hours}$.
   b. $143\ \text{mg}$ of medicine four times a day (every $6\ \text{hours}$).

34. $259\ \text{mg}$ (rounding to 2 significant figures); toxic dose for this weight is $4310\ \text{mg}\,(4.3\ \text{g})$.

35. $41\ ^\circ\text{C}$ and $314\ \text{K}$

36. $102\ ^\circ\text{F};$ this is below the $(105\ ^\circ F)$ to be considered heat stroke.