TIMBERLAKE

CHEMISTRY

An Introduction to General, Organic, and Biological Chemistry

ELEVENTH EDITION

Periodic Table of Elements

Representative elements

Alkali metals — Alkaline earth metals — Halogens — Noble gases

Transition elements

Period number	1 Group 1A	2 Group 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Group 3A	14 Group 4A	15 Group 5A	16 Group 6A	17 Group 7A	18 Group 8A
1	1 H 1.008																	2 He 4.003
2	3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.41	31 Ga 69.72	32 Ge 72.64	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (99)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57* La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89† Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (265)	109 Mt (266)	110 Ds (271)	111 Rg (272)	112 Cn (285)	113 — (284)	114 — (289)	115 — (288)	116 — (292)		118 — (294)

*Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
†Actinides	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Metals — Metalloids — Nonmetals

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Chemistry and Measurements

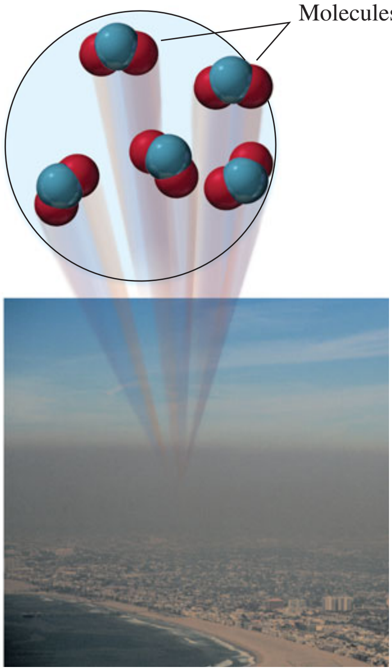
1

LOOKING AHEAD

"I use measurement in just about every part of my nursing practice," says registered nurse Vicki Miller. "When I receive a doctor's order for a medication, I have to verify that order. Then I draw a carefully measured volume from an IV or a vial to create that particular dose. Some dosage orders are specific to the size of the patient. I measure the patient's weight and calculate the dosage required for the weight of that patient."

Nurses use measurement each time they take a patient's temperature, height, weight, or blood pressure. Measurement is used to obtain the correct amounts for injections and medications and to determine the volumes of fluid intake and output. For each measurement, the amounts and units are recorded in the patient's records.

What are some questions in chemistry you have been curious about? Perhaps you are interested in how smog is formed, what causes ozone depletion, or how aspirin relieves a headache. Just like you, chemists are curious about the world we live in.



The chemical reaction of NO with oxygen in the air forms NO_2, which produces the reddish brown color of smog.

- How does car exhaust produce the smog that hangs over our cities? One component of car exhaust is nitrogen oxide (NO) formed in car engines, where high temperatures convert nitrogen gas (N_2) and oxygen gas (O_2) to NO gas. In chemistry, these reactions are written in the form of equations such as $N_2(g) + O_2(g) \longrightarrow 2NO(g)$. The reaction of NO with oxygen in the air produces NO_2, which gives the reddish brown color to smog.
- Why has the ozone layer been depleted in certain parts of the atmosphere? In the 1970s, scientists associated substances called chlorofluorocarbons (CFCs) with the depletion of ozone over Antarctica. As CFCs are broken down by ultraviolet (UV) light, chlorine (Cl) is released that causes the breakdown of ozone (O_3) molecules and destroys the ozone layer:

$$Cl + O_3 \longrightarrow ClO + O_2$$

- Why does aspirin relieve a headache? When a part of the body is injured, substances called prostaglandins are produced, which cause inflammation and pain. Aspirin acts to block the production of prostaglandins, thereby reducing inflammation, pain, and fever.

Chemists assess the impact of chemistry on our lives and our environment by making measurements. Levels of toxic materials in the air, soil, and water are measured. Measurements help us understand about radon in our homes, global warming, trans fatty acids, and DNA analysis. Understanding chemistry and measurement helps us make proper choices about our world.

Think about your day; you probably made some measurements. Perhaps you checked your weight by stepping on a scale. If you cooked some rice for dinner, you added 2 cups of water to 1 cup of rice. If you did not feel well, you may have taken your temperature. If you stopped at the gas station, you watched the gas pump measure the number of gallons of gasoline you put in the car.

Your weight on a bathroom scale is a measurement.

Measurement is an essential part of health careers such as nursing, dental hygiene, respiratory therapy, nutrition, and veterinary technology. The temperature, height, and weight of a patient are measured and recorded. Samples of blood and urine are collected and sent to a laboratory where glucose, pH, urea, and protein are measured by the clinical technicians. By learning about measurement, you develop skills for solving problems and learn how to work with numbers.

1.1 Chemistry and Chemicals

LEARNING GOAL

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

Chemistry is the study of the composition, structure, properties, and reactions of matter. *Matter* is another word for all the substances that make up our world. Perhaps you imagine that chemistry is done only in a laboratory by a chemist wearing a lab coat and goggles. Actually, chemistry happens all around you every day and has a big impact on everything you use and do. You are doing chemistry when you cook food, add chlorine to a swimming pool, or drop an antacid tablet into water. Plants grow because chemical reactions convert carbon dioxide, water, and energy to carbohydrates. Chemical reactions take place when you digest food and break it down into substances that you need for energy and health.

Antacid tablets undergo a chemical reaction when dropped into water.

Branches of Chemistry

The field of chemistry is divided into branches such as organic, inorganic, and general chemistry. Organic chemistry is the study of substances that contain the element carbon. Inorganic chemistry is the study of all other substances except those that contain carbon. General chemistry is the study of the composition, properties, and reactions of matter.

Today, chemistry is often combined with other sciences such as geology, biology, and physics to form cross disciplines such as geochemistry, biochemistry, and physical chemistry. Geochemistry is the study of the chemical composition of ores, soils, and minerals of the surface of Earth and other planets. Biochemistry is the study of the chemical reactions that take place in biological systems. Physical chemistry is the study of the physical nature of chemical systems including energy changes.

Biochemists analyze laboratory samples.

Chemistry Link to History

EARLY CHEMISTS: THE ALCHEMISTS

For many centuries, chemists have studied changes in various substances. From the time of the ancient Greeks to about the sixteenth century, alchemists described a substance in terms of four components of nature: earth, air, fire, and water. By the eighth century, alchemists searched for an unknown substance called a philosopher's stone that they thought would turn metals into gold as well as prolong youth and postpone death. Although these efforts failed, the alchemists did provide information on the processes and chemical reactions involved in the extraction of metals from ores. The alchemists also designed some of the first laboratory equipment and developed early laboratory procedures.

The alchemist Paracelsus (1493–1541) thought that alchemy should be about preparing new medicines, not about producing gold. Using observation and experimentation, he proposed that a healthy body was regulated by a series of chemical processes that could be unbalanced by certain chemical compounds and rebalanced by using minerals and medicines. For example, he determined that inhaled dust, not underground spirits, caused lung disease in miners. He also thought that goiter was a problem caused by contaminated water, and he treated syphilis with compounds of mercury. His opinion of medicines was that the right dose makes the difference between a poison and a cure. Today, this idea is part of the risk analysis of medicines. Paracelsus changed alchemy in ways that helped to establish modern medicine and chemistry.

Alchemists in the Middle Ages developed laboratory procedures.

Swiss alchemist and physician Paracelsus (1493–1541) believed that chemicals could be used as medicines.

TABLE 1.1 Chemicals Commonly Used in Toothpaste

Chemical	Function
Calcium carbonate	An abrasive used to remove plaque
Sorbitol	Prevents loss of water and hardening of toothpaste
Glycerin	Makes toothpaste foam in the mouth
Sodium lauryl sulfate	A detergent used to loosen plaque
Titanium dioxide	Makes toothpaste white and opaque
Triclosan	Inhibits bacteria that cause plaque and gum disease
Sodium fluorophosphate	Prevents formation of cavities by strengthening tooth enamel with fluoride
Methyl salicylate	Gives toothpaste a pleasant wintergreen flavor

Chemicals

Toothpaste is a combination of many chemicals.

All the things you see around you are composed of one or more chemicals. A **chemical** is a substance that always has the same composition and properties wherever it is found. Chemical processes take place in chemistry laboratories, manufacturing plants, and pharmaceutical labs as well as every day in nature and in our bodies. Often the terms *chemical* and *substance* are used interchangeably to describe a specific type of material.

Every day you use products containing substances that were prepared by chemists. Soaps and shampoos contain chemicals that remove oils on your skin and scalp. When you brush your teeth, the chemicals in toothpaste clean your teeth, prevent plaque formation, and stop tooth decay. Some of the substances used to make toothpaste are listed in Table 1.1.

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

FIGURE 1.1 Many of the items found in a kitchen are obtained using chemical reactions.

Q What are some other chemicals found in a kitchen?

SAMPLE PROBLEM 1.1

Chemicals

Why is the copper in copper wire an example of a chemical, while sunlight is not?

SOLUTION

Copper is a chemical that has the same composition and properties wherever it is found. Sunlight is energy given off by the Sun. Thus, sunlight does not contain matter, which means it is not a chemical.

STUDY CHECK 1.1

Which of the following are chemicals?

a. iron **b.** tin
c. a low temperature **d.** water

The answers to all the Study Checks can be found at the end of this chapter.

QUESTIONS AND PROBLEMS

Chemistry and Chemicals

In every chapter, each magenta, odd-numbered exercise in *Questions and Problems* is paired with the next even-numbered exercise. The answers for the magenta, odd-numbered Questions and Problems are given at the end of this chapter. The complete solutions to the odd-numbered Questions and Problems are in the *Study Guide*.

1.1 Obtain a bottle of multivitamins, and observe the list of ingredients. List four. Which ones are chemicals?

1.2 Obtain a box of breakfast cereal, and observe the list of ingredients. List four. Which ones are chemicals?

1.3 A “chemical-free” shampoo includes the ingredients water, cocomide, glycerin, and citric acid. Is the shampoo truly “chemical-free”?

1.4 A “chemical-free” sunscreen includes the ingredients titanium dioxide, vitamin E, and vitamin C. Is the sunscreen truly “chemical-free”?

1.2 A Study Plan for Learning Chemistry

LEARNING GOAL

Develop a study plan for learning chemistry.

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

Features in This Text Help You Study Chemistry

This text has been designed with a variety of study aids to complement your individual learning style. On the inside of the front cover is a periodic table of the elements. On the inside of the back cover are tables that summarize useful information needed throughout your study of chemistry. Each chapter begins with Looking Ahead, which outlines the topics in the chapter. A Learning Goal at the beginning of each section previews the concepts you are to learn. At the end of the text, there is a comprehensive Glossary and Index, which list and define key terms used in the text.

Before you begin reading, obtain an overview of a chapter by reviewing the topics in Looking Ahead. As you prepare to read a section of the chapter, look at the section title and turn it into a question. For example, for Section 1.1 “Chemistry and Chemicals,” you could ask “What is chemistry?” or “What are chemicals?” When you are ready to read through that section, review the Learning Goal, which tells you what to expect in that section. As you read, try to answer your question. Throughout the chapter, you will find Concept Checks that will help you understand key ideas. When you come to a Sample Problem, take the time to work it through and try the associated Study Check. Many Sample Problems are accompanied by a visual Guide to Problem Solving (GPS). At the end of each section, you will find a set of Questions and Problems that allows you to apply problem solving immediately to the new concepts.

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

Studying in a group can be beneficial to learning.

At the end of each chapter, you will find several study aids that complete the chapter. *Concept Maps* show the connections between important concepts and *Chapter Reviews* provide a summary. The *Key Terms* are in boldface type in the text and listed with their definitions. *Understanding the Concepts*, a set of questions that uses art and structures, helps you visualize concepts. *Additional Questions and Problems* and *Challenge Problems* provide additional problems to test your understanding of the topics in the chapter. All the problems are paired, and the answers to all the *Study Checks* and to the odd-numbered *Questions and Problems* are provided. If your answers match, you most likely understand the topic; if not, you need to study the section again.

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

Using Active Learning to Learn Chemistry

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

Students discuss a chemistry problem with their professor during office hours.

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

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

Thinking about Your Study Plan

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

TABLE 1.2 Steps in Active Learning

1. Read each *Learning Goal* for an overview of the material.
2. Form a question from the title of the section you are going to read.
3. Read the section looking for answers to your question.
4. Self-test by working *Concept Checks*, *Sample Problems*, and *Study Checks*.
5. Complete the *Questions and Problems* that follow that section, and check the answers for the magenta odd-numbered problems.
6. Work the exercises in the *Study Guide* and go to *www.masteringchemistry.com* for self-study materials and instructor-assigned homework (optional).
7. Proceed to the next section, and repeat the previous steps.

My study of chemistry will include the following:

_______ reading the chapter before a lecture
_______ going to lectures
_______ reviewing the *Learning Goals*
_______ keeping a problem notebook
_______ reading the text as an active learner
_______ self-testing by working *Questions and Problems* following each section and checking solutions in the text
_______ being an active learner during lectures
_______ organizing a study group
_______ seeing the professor during office hours
_______ completing exercises in the *Study Guide*
_______ working through the tutorials at *www.masteringchemistry.com*
_______ attending review sessions
_______ organizing my own review sessions
_______ studying in small doses as often as I can

CONCEPT CHECK 1.1

A Study Plan for Learning Chemistry

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

a. skipping lectures
b. forming a study group
c. keeping a problem notebook
d. waiting to study the night before the exam
e. becoming an active learner

ANSWER

Your success in chemistry can be improved by
b. forming a study group
c. keeping a problem notebook
e. becoming an active learner

QUESTIONS AND PROBLEMS

A Study Plan for Learning Chemistry

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

1.6 What are four things that would make it difficult for you to succeed in chemistry?

1.7 A student in your class asks you for advice on learning chemistry. Which of the following might you suggest?
a. Form a study group.
b. Skip lectures.
c. Visit the professor during office hours.
d. Wait until the night before an exam to study.
e. Become an active learner.
f. Work the exercises in the *Study Guide*.

1.8 A student in your class asks you for advice on learning chemistry. Which of the following might you suggest?
a. Do the assigned problems.
b. Don't read the book; it's never on the test.
c. Attend review sessions.
d. Read the assignment before a lecture.
e. Keep a problem notebook.
f. Do the tutorials at *www.masteringchemistry.com*.

LEARNING GOAL

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

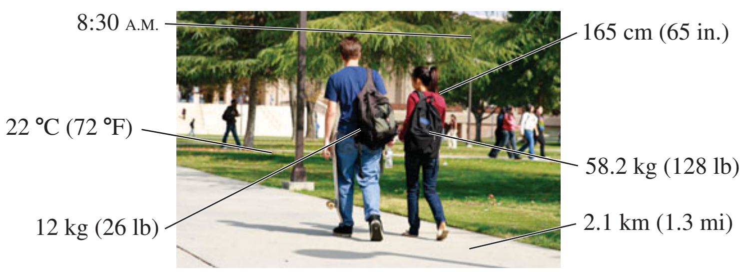
1.3 Units of Measurement

Scientists and health professionals throughout the world use the **metric system** of measurement. It is also the common measuring system in all but a few countries in the world. In 1960, scientists adopted a modification of the metric system called the **International System of Units (SI)**, or Système International, which is now the official system of measurement throughout the world except for the United States. In chemistry, we use metric units and SI units for length, volume, mass, temperature, and time, as listed in Table 1.3.

TABLE 1.3 Units of Measurement

Measurement	Metric	SI
Length	meter (m)	meter (m)
Volume	liter (L)	cubic meter (m^3)
Mass	gram (g)	kilogram (kg)
Temperature	degree Celsius (°C)	kelvin (K)
Time	second (s)	second (s)

Suppose today you walk 2.1 km to campus carrying a backpack that has a mass of 12 kg. Perhaps it is 8:30 A.M. and the temperature is 22 °C. You have a mass of 58.2 kg and a height of 165 cm. You may be more familiar with these measurements stated in the U.S. system of measurement; then you would walk 1.3 mi carrying a backpack that weighs 26 lb. The temperature at 8:30 A.M. would be 72 °F. You have a weight of 128 lb and a height of 65 in.



There are many measurements in everyday life.

Length

The metric and SI unit of length is the **meter (m)**. A meter is 39.4 inches (in.), which makes it slightly longer than a yard (yd). The **centimeter (cm)**, a smaller unit of length, is commonly used in chemistry and is about equal to the width of your little finger. For comparison, there are 2.54 cm in 1 in. (see Figure 1.2).

Some relationships between units for length are

$$1\ \text{m} = 100\ \text{cm}$$
$$1\ \text{m} = 39.4\ \text{in.}$$
$$1\ \text{m} = 1.09\ \text{yd}$$
$$2.54\ \text{cm} = 1\ \text{in.}$$

Volume

Volume is the amount of space a substance occupies. A **liter (L)** is slightly larger than the quart (qt), (1 L = 1.06 qt). In a laboratory or a hospital, chemists work with metric units

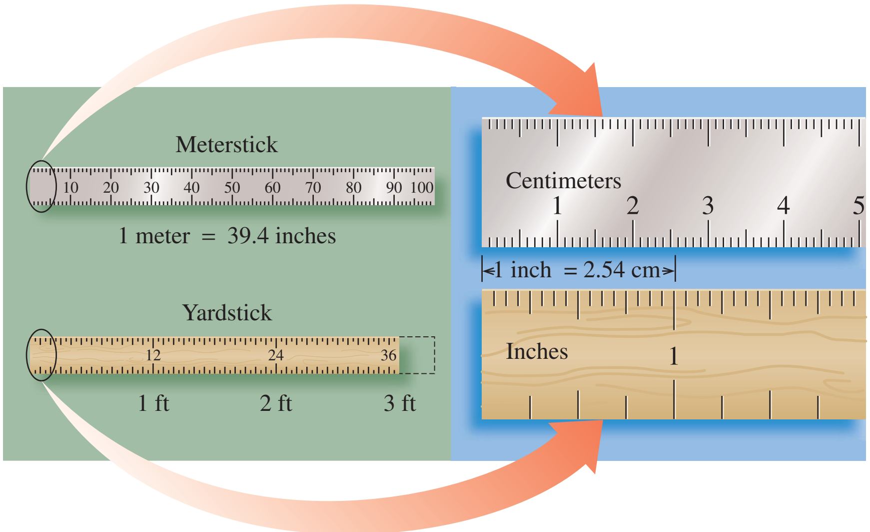


FIGURE 1.2 Length in the metric (SI) system is based on the meter, which is slightly longer than a yard.
Q How many centimeters are in a length of 1 inch?

of volume that are smaller and more convenient, such as the **milliliter (mL)**. There are 1000 mL in 1 L. A comparison of metric and U.S. units for volume appears in Figure 1.3.

Some relationships between units for volume are

$$1\text{ L} = 1000\text{ mL}$$
$$1\text{ L} = 1.06\text{ qt}$$
$$946\text{ mL} = 1\text{ qt}$$

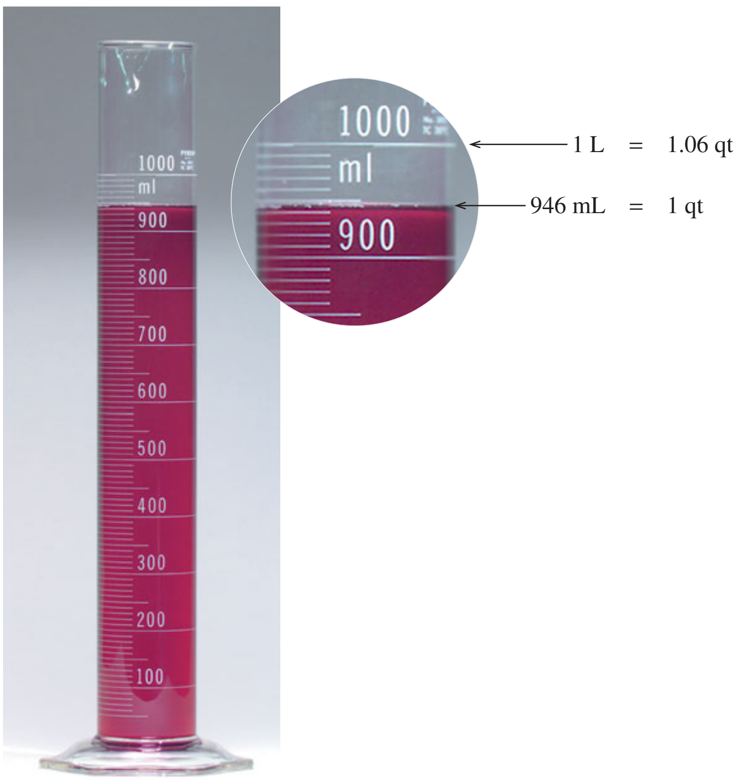


FIGURE 1.3 Volume is the space occupied by a substance. In the metric system, volume is based on the liter, which is slightly larger than a quart.
Q How many milliliters are in 1 quart?

Mass

The **mass** of an object is a measure of the quantity of material it contains. The SI unit of mass, the **kilogram (kg)**, is used for larger masses such as body mass. In the metric system, the unit for mass is the **gram (g)**, which is used for smaller masses. There are 1000 g in one kilogram. It takes 2.20 lb to make 1 kg, and 454 g is equal to 1 lb.

The standard kilogram for the United States is stored at the National Institute of Standards and Technology (NIST).

FIGURE 1.4 On an electronic balance, a nickel has a mass of 5.01 g in the digital readout.
Q What is the mass of 10 nickels?

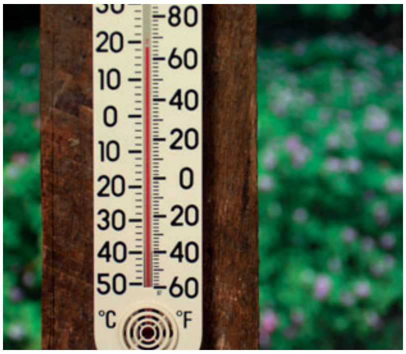


FIGURE 1.5 A thermometer is used to determine temperature.
Q What kinds of temperature readings have you made today?

A stopwatch is used to measure the time of a race.

Some relationships between units for mass are

$1 \text{ kg} = 1000 \text{ g}$

$1 \text{ kg} = 2.20 \text{ lb}$

$454 \text{ g} = 1 \text{ lb}$

You may be more familiar with the term *weight* than with mass. Weight is a measure of the gravitational pull on an object. On Earth, an astronaut with a mass of 75.0 kg has a weight of 165 lb. On the moon where the gravitational pull is one-sixth that of Earth, the astronaut has a weight of 27.5 lb. However, the mass of the astronaut is the same as on Earth, 75.0 kg. Scientists measure mass rather than weight because mass does not depend on gravity.

In a chemistry laboratory, an analytical balance is used to measure the mass in grams of a substance (see Figure 1.4).

Temperature

Temperature tells us how hot something is, how cold it is outside, or helps us determine if we have a fever (see Figure 1.5). In the metric system, temperature is measured using Celsius temperature. On the **Celsius (°C) temperature scale**, water freezes at 0 °C and boils at 100 °C, while on the Fahrenheit (°F) scale, water freezes at 32 °F and boils at 212 °F. In the SI system, temperature is measured using the **Kelvin (K) temperature scale** on which the lowest possible temperature is 0 K. A unit on the Kelvin scale is called a kelvin (K) and is not written with a degree sign. We will discuss the relationships between these three temperature scales in Chapter 2.

Time

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

CONCEPT CHECK 1.2

Units of Measurement

State the type of measurement indicated by each of the following metric units:

a. gram **b.** liter
c. centimeter **d.** degree Celsius

ANSWER

a. A gram is a unit of mass.
b. A liter is a unit of volume.
c. A centimeter is a unit of length.
d. A degree Celsius is a unit of temperature.

Explore Your World

UNITS LISTED ON LABELS

Read the labels on a variety of products such as sugar, salt, soft drinks, vitamins, and dental floss.

QUESTIONS

1. What metric or SI units of measurement are listed on the labels?
2. What type of measurement (mass, volume, etc.) do they indicate?
3. Write the metric or SI amounts in terms of a number plus a unit.

QUESTIONS AND PROBLEMS

Units of Measurement

1.9 Compare the units you would use and the units that a student in Mexico would use to measure each of the following:
a. your body mass
b. your height
c. amount of gasoline to fill a gas tank
d. temperature at noon

1.10 Why are each of the following statements confusing, and how would you make them clear using metric (SI) units?
a. I rode my bicycle for a distance of 15 today.
b. My dog weighs 15.
c. It is hot today. It is 30.
d. I lost 1.5 last week.

1.11 State the name of the unit and the type of measurement indicated for each of the following quantities:
a. 4.8 m **b.** 325 g **c.** 1.5 mL
d. 480 s **e.** 28 °C

1.12 State the name of the unit and the type of measurement indicated for each of the following quantities:
a. 0.8 L **b.** 3.6 cm **c.** 4 kg
d. 35 lb **e.** 373 K

1.4 Scientific Notation

LEARNING GOAL
Write a number in scientific notation.

MC TUTORIAL
Scientific Notation

Scientific Notation

In chemistry, we use numbers that are very large or very small. We might measure something as tiny as the width of a human hair, which is about 0.000 008 m. Or perhaps we want to count the number of hairs in the average human scalp, which is about 100 000 hairs (see Figure 1.6). In this text, we add spaces between sets of three digits when it helps make the places easier to count. However, it is more convenient to write large and small numbers in *scientific notation*.

Item	Standard Number	Scientific Notation
Width of a human hair	0.000 008 m	8×10^{-6} m
Hairs on a human scalp	100 000 hairs	1×10^{5} hairs

FIGURE 1.6 Humans have an average of 1×10^5 hairs on their scalps. Each hair is about 8×10^{-6} m wide.
Q Why are large and small numbers written in scientific notation?

Writing a Number in Scientific Notation

A number written in **scientific notation** has three parts: a coefficient, a power of 10, and a measurement unit. For example, the number 2400 m is written in scientific notation as 2.4×10^3 m. The coefficient is 2.4, the 3 is the power of 10, and m is the measurement unit of meters. The coefficient is determined by moving the decimal point to the left to give a coefficient that is at least 1 but less than 10. Because we moved the decimal point three places, the power of 10 is 3, which is written as 10^3. For any number greater than 1, the power of 10 is positive.

$$\underset{\leftarrow 3\ \text{places}}{2400.}\ \text{m} = 2.4 \times 1000 = \underbrace{2.4}_{\text{Coefficient}} \times \underbrace{10^3}_{\text{Power of 10}}\ \text{m}$$

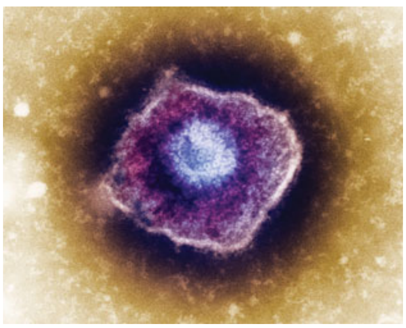
When a number less than 1 is written in scientific notation, the power of 10 is a negative number. For example, the number 0.000 86 g is written in scientific notation by moving the decimal point four places to give a coefficient of 8.6. Because the decimal point was moved four places to the right, the power of 10 is a negative 4, written as 10^{-4}.

$$\underset{4\ \text{places}\rightarrow}{0.00086}\ \text{g} = \frac{8.6}{10\,000} = \frac{8.6}{10 \times 10 \times 10 \times 10} = \underbrace{8.6}_{\text{Coefficient}} \times \underbrace{10^{-4}}_{\text{Power of 10}}\ \text{g}$$

Table 1.4 gives some examples of numbers written as positive and negative powers of 10. The powers of 10 are really a way of keeping track of the decimal point in the decimal number. Table 1.5 gives several examples of writing measurements in scientific notation.

TABLE 1.4 Some Powers of 10

Number	Multiples of 10	Scientific Notation	
10 000	$10 \times 10 \times 10 \times 10$	1×10^4	
1 000	$10 \times 10 \times 10$	1×10^3	Some positive powers of 10
100	10×10	1×10^2	
10	10	1×10^1	
1	0	1×10^0	
0.1	$\frac{1}{10}$	1×10^{-1}	
0.01	$\frac{1}{10} \times \frac{1}{10} = \frac{1}{100}$	1×10^{-2}	Some negative powers of 10
0.001	$\frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} = \frac{1}{1\,000}$	1×10^{-3}	
0.0001	$\frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} = \frac{1}{10\,000}$	1×10^{-4}	



A chickenpox virus has a diameter of 3×10^{-7} m.

TABLE 1.5 Some Measurements Written in Scientific Notation

Measured Quantity	Standard Number	Scientific Notation
Volume of gasoline used in United States each year	550 000 000 000 L	5.5×10^{11} L
Diameter of Earth	12 800 000 m	1.28×10^7 m
Time for light to travel from the Sun to Earth	500 s	5×10^2 s
Mass of a typical human	68 kg	6.8×10^1 kg
Mass of a hummingbird	0.002 kg	2×10^{-3} kg
Diameter of a chickenpox (varicella zoster) virus	0.000 000 3 m	3×10^{-7} m
Mass of bacterium (mycoplasma)	0.000 000 000 000 000 000 1 kg	1×10^{-19} kg

Scientific Notation and Calculators

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

Number to Enter	Procedure	Calculator Display
4×10^6	4 EXP (EE) 6	4 06 or 4^{06} or 4 E06
2.5×10^{-4}	2.5 EXP (EE) +/− 4	2.5−04 or 2.5^{-04} or 2.5 E−04

When a calculator display appears in scientific notation, it is shown as a number between 1 and 10 followed by a space and the power of 10. To express this display in scientific notation, write the coefficient value, write $\times$ 10, and use the power of 10 as an exponent.

Calculator Display	Expressed in Scientific Notation
7.52 04 or 7.52^{04} or 7.52 E04	7.52×10^4
5.8−02 or 5.8^{-02} or 5.8 E−02	5.8×10^{-2}

SAMPLE PROBLEM 1.2

Scientific Notation

Write each of the following in scientific notation:

a. 75 000 m **b.** 0.0098 g **c.** 143 mL

SOLUTION

a. To write a coefficient of 7.5, which is more than 1 but less than 10, move the decimal point four places to the left to give 7.5×10^4 m.

b. To write a coefficient of 9.8, which is more than 1 but less than 10, move the decimal point three places to the right to give 9.8×10^{-3} g.

c. To write a coefficient of 1.43, which is more than 1 but less than 10, move the decimal point two places to the left to give 1.43×10^2 mL.

STUDY CHECK 1.2

Write each of the following in scientific notation:

a. 425 000 m **b.** 0.000 000 86 g

QUESTIONS AND PROBLEMS

Scientific Notation

1.13 Write each of the following in scientific notation:
a. 55 000 m **b.** 480 g **c.** 0.000 005 cm
d. 0.000 14 s **e.** 0.0072 L **f.** 670 000 kg

1.14 Write each of the following in scientific notation:
a. 180 000 000 g **b.** 0.000 06 m **c.** 750 °C
d. 0.15 mL **e.** 0.024 s **f.** 1500 cm

1.15 Which number in each of the following pairs is larger?
a. 7.2×10^3 cm or 8.2×10^2 cm
b. 4.5×10^{-4} kg or 3.2×10^{-2} kg
c. 1×10^4 L or 1×10^{-4} L
d. 0.000 52 m or 6.8×10^{-2} m

1.16 Which number in each of the following pairs is smaller?
a. 4.9×10^{-3} s or 5.5×10^{-9} s
b. 1250 kg or 3.4×10^2 kg
c. 0.000 000 4 m or 5.0×10^2 m
d. 2.50×10^2 g or 4×10^5 g

LEARNING GOAL

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

1.5 Measured Numbers and Significant Figures

When you make a measurement, you use some type of measuring device. For example, you may use a meterstick to measure your height, a scale to check your weight, or a thermometer to take your temperature. **Measured numbers** are the numbers you obtain when you measure a quantity such as your height, weight, or temperature.

Measured Numbers

Suppose you are going to measure the lengths of the objects in Figure 1.7. You would select a metric ruler that may have lines marked in 1 cm divisions or perhaps in divisions of 0.1 cm. To report the length of the object, you observe the numerical values of the marked lines at the end of object. Then, you can *estimate* by visually dividing the space between the marked lines. This estimated value is the final digit in a measured number.

For example, in Figure 1.7a, the end of the object is between the marks of 4 cm and 5 cm, which means that the length is more than 4 cm but less than 5 cm. This is written as 4 cm plus an estimated digit. If you estimate that the end of the object is halfway between 4 cm and 5 cm, you would report its length as 4.5 cm. The last digit in a measured number may differ because people do not always estimate in the same way. Thus, someone else might report the length of the same object as 4.4 cm.

The metric ruler shown in Figure 1.7b is marked with lines at every 0.1 cm. With this ruler, you can estimate to the hundredths place (0.01 cm). Now you can determine that the end of the object is between 4.5 cm and 4.6 cm. Perhaps you report the length of the object as 4.55 cm, while another student reports its length as 4.56 cm. Both results are acceptable.

In Figure 1.7c, the end of the object appears to line up with the 3-cm mark. Because the divisions are marked in units of 1 cm, the length of the object is between 3 cm and 4 cm. Because the end of the object is on the 3-cm mark, the estimated digit is 0, which means the measurement is reported as 3.0 cm.

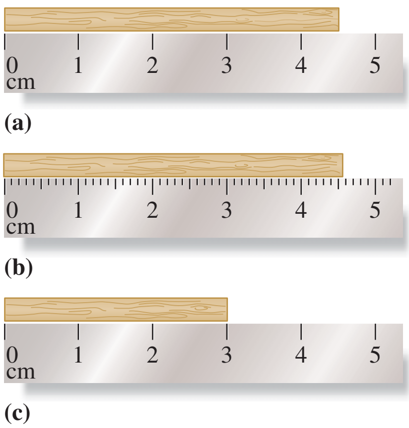


FIGURE 1.7 The lengths of the rectangular objects are measured as **(a)** 4.5 cm and **(b)** 4.55 cm.
Q What is the length of the object in (c)?

SELF STUDY ACTIVITY Significant Figures

TUTORIAL Counting Significant Figures

Significant Figures

In a measured number, the **significant figures (SFs)** are all the digits including the estimated digit. Nonzero numbers are always counted as significant figures. However, a zero may or may not be a significant figure depending on its position in a number. Table 1.6 gives the rules and examples of counting significant figures.

TABLE 1.6 Significant Figures in Measured Numbers

Rule	Measured Number	Number of Significant Figures
1. A number is a *significant figure* if it is		
a. not a zero	4.5 g 122.35 m	2 5
b. a zero between nonzero digits	205 m 5.082 kg	3 4
c. a zero at the end of a decimal number	50. L 25.0 °C 16.00 g	2 3 4
d. in the coefficient of a number written in scientific notation	4.8×10^5 m 5.70×10^{-3} g	2 3
2. A zero is *not significant* if it is		
a. at the beginning of a decimal number	0.0004 s 0.075 cm	1 2
b. used as a placeholder in a large number without a decimal point	850 000 m 1 250 000 g	2 3

When one or more zeros in a large number are significant digits, they are shown more clearly by writing the number using scientific notation. For example, if the first zero in the measurement 500 m is significant, it is written as 5.0×10^2 m. In this text, we will place a decimal point at the end of a number if the zeros are significant. For example, a measurement written as 500. g indicates that all the zeros are significant. Thus 500. g has three significant figures. It could also be written as 5.00×10^2 g. We will assume that zeros at the end of large numbers without a decimal point are not significant. Therefore, we would write a value of 500 g as 5×10^2 g, which has only one significant figure.

CONCEPT CHECK 1.3

Significant Zeros

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

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

ANSWER

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

Exact Numbers

Exact numbers are numbers obtained by counting items or using a definition that compares two units in the same measuring system. Suppose a friend asks you how many classes you are taking. You would answer by counting the number of classes in your schedule. It would not use any measuring tool. Suppose you are asked to state the number of seconds in one minute. Without using any measuring device, you would give the definition: There are 60 seconds in 1 minute. Exact numbers are not measured, do not have a limited number of significant figures, and do not affect the number of significant figures in a calculated answer. For more examples of exact numbers, see Table 1.7.

The number of baseballs is counted, which means 2 is an exact number.

TABLE 1.7 Examples of Some Exact Numbers

	Defined Equalities	
Counted Numbers	**U.S. System**	**Metric System**
8 doughnuts	1 ft = 12 in.	1 L = 1000 mL
2 baseballs	1 qt = 4 cups	1 m = 100 cm
5 capsules	1 lb = 16 oz	1 kg = 1000 g

CONCEPT CHECK 1.4

Measured Numbers and Significant Figures

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

a. 42.2 g **b.** three eggs **c.** 5.0×10^{-3} cm
d. 450 000 km **e.** The value of 12 in. for 1 ft = 12 in.

ANSWER

a. The value 42.2 g is a measured number, which has three SFs because all nonzero digits are significant.
b. The value three eggs is obtained by counting, which makes it an exact number.
c. The value 5.0×10^{-3} cm is a measured number, which has two SFs because all numbers in the coefficient of a number written in scientific notation are significant.
d. The value 450 000 km is a measured number, which has two SFs because the zeros in a large number without a decimal point are not significant.
e. The value 12 in. is exact because the relationship 1 ft = 12 in. is a definition in the U.S. system of measurement.

QUESTIONS AND PROBLEMS

Measured Numbers and Significant Figures

1.17 Identify the numbers in each of the following statements as measured or exact:
a. A patient weighs 155 lb.
b. A patient is given 2 tablets of medication.
c. In the metric system, 1 kg is equal to 1000 g.
d. The distance from Denver, Colorado, to Houston, Texas, is 1720 km.

1.18 Identify the numbers in each of the following statements as measured or exact:
a. There are 31 students in the laboratory.
b. The oldest known flower lived 1.20×10^8 years ago.
c. The largest gem ever found, an aquamarine, has a mass of 104 kg.
d. A laboratory test shows a blood cholesterol level of 184 mg/dL.

1.19 Identify the measured number(s), if any, in each of the following pairs of numbers:
a. 3 hamburgers and 6 oz of hamburger
b. 1 table and 4 chairs
c. 0.75 lb of grapes and 350 g of butter
d. 60 s = 1 min

1.20 Identify the exact number(s), if any, in each of the following pairs of numbers:
a. 5 pizzas and 50.0 g of cheese
b. 6 nickels and 16 g of nickel
c. 3 onions and 3 lb of onions
d. 5 miles and 5 cars

1.21 Indicate if the zeros are significant or not in each of the following measurements:
a. 0.0038 m **b.** 5.04 cm **c.** 800. L
d. 3.0×10^{-3} kg **e.** 85 000 g

1.22 Indicate if the zeros are significant or not in each of the following measurements:
a. 20.05 °C **b.** 5.00 m **c.** 0.000 02 g
d. 120 000 years **e.** 8.05×10^2 L

1.23 How many significant figures are in each of the following measured quantities?
a. 11.005 g **b.** 0.000 32 m **c.** 36 000 000 km
d. 1.80×10^4 kg **e.** 0.8250 L **f.** 30.0 °C

1.24 How many significant figures are in each of the following measured quantities?
a. 20.60 mL **b.** 1036.48 kg **c.** 4.00 m
d. 20.8 °C **e.** 60 800 000 g **f.** 5.0×10^{-3} L

1.25 In which of the following pairs do both numbers contain the same number of significant figures?
a. 11.0 m and 11.00 m
b. 405 K and 504.0 K
c. 0.000 12 s and 12 000 s
d. 250.0 L and 2.5×10^{-2} L

1.26 In which of the following pairs do both numbers contain the same number of significant figures?
a. 0.005 75 g and 5.75×10^{-3} g
b. 0.0250 m and 0.205 m
c. 150 000 s and 1.50×10^4 s
d. 3.8×10^{-2} L and 7.5×10^5 L

1.6 Significant Figures in Calculations

LEARNING GOAL

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

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

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

TUTORIAL
Significant Figures in Calculations

Rounding Off

Suppose you decide to buy carpeting for a room that measures 5.52 m by 3.58 m. Each of these measurements has three significant figures because the measuring tape limits your estimated place to 0.01 m. To determine how much carpeting you need, you would calculate the area of the room by multiplying 5.52 times 3.58 on your calculator. The calculator shows the numbers 19.7616 in its display. However, this is not the correct final answer because there are too many numbers, which is the result of the multiplication process. Because each of the original measurements has only three significant figures, the numbers 19.7616 in the calculator display must be *rounded off* to obtain an answer that also has three significant figures, which is 19.8. Therefore, you can order carpeting that will cover an area of 19.8 m^2.

A technician uses a calculator in the laboratory.

Each time you use a calculator, it is important to look at the measurements and determine the number of significant figures that can be used for the answer. You can use the following rules to round off the numbers shown in a calculator display.

TUTORIAL
Using Significant Figures

Rules for Rounding Off

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

	Three Significant Figures	Two Significant Figures
Example 1: 8.4234 rounds off to	8.42	8.4
Example 2: 14.780 rounds off to	14.8	15
Example 3: 3256 rounds off to	3260* (3.26×10^3)	3300* (3.3×10^3)

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

CONCEPT CHECK 1.5

Rounding Off

Select the correct value when 2.8456 m is rounded off to each of the following:

a. three significant figures:	2.84 m	2.85 m	2.8 m	2.90 m
b. two significant figures:	2.80 m	2.85 m	2.8 m	2.90 m

ANSWER

a. To round off 2.8456 m to three significant figures, drop the final digits 56 and increase the last retained digit by 1 to give 2.85 m.

b. To round off 2.8456 m to two significant figures, drop the final digits 456 to give 2.8 m.

SAMPLE PROBLEM 1.3

Rounding Off

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

a. 35.7823 m
b. 0.002 625 L
c. 3.8268×10^3 g
d. 1.2836 kg

SOLUTION

a. 35.8 m
b. 0.002 63 L
c. 3.83×10^3 g
d. 1.28 kg

STUDY CHECK 1.3

Round off each of the numbers in Sample Problem 1.3 to two significant figures.

A calculator is helpful in working problems and doing calculations faster.

Multiplication and Division

In multiplication or division, the final answer is written so it has the same number of significant figures as the measurement with the *fewest significant figures* (SFs).

Example 1

Multiply the following measured numbers: 24.65×0.67

24.65	$\boxed{\times}$	0.67	$\boxed{=}$	16.5155	$\longrightarrow$	17
Four SFs		Two SFs		Calculator display		Final answer, rounded off to two SFs

The answer in the calculator display has more digits than the data allow. The measurement 0.67 has the fewer number of significant figures, two. Therefore, the calculator answer is rounded off to two significant figures.

Example 2

Solve the following:

$$\frac{2.85 \times 67.4}{4.39}$$

To do this problem on a calculator, we might press the keys in the following order:

2.85	$\boxed{\times}$	67.4	$\boxed{\div}$	4.39	$\boxed{=}$	43.756264	$\longrightarrow$	43.8
Three SFs		Three SFs		Three SFs		Calculator display		Final answer, rounded off to three SFs

All the measurements in this problem have three significant figures. Therefore, the calculator result is rounded off to give an answer, 43.8, that has three significant figures.

Adding Significant Zeros

Sometimes, a calculator display gives a small whole number. To report an answer with the correct number of significant figures, you may need to write significant zeros after the calculator numbers. For example, suppose the calculator display is 4, but the measurements each have three significant figures. Then the display of 4 is written as 4.00 by adding two significant zeros.

$\frac{8.00}{2.00}$	=	4	$\longrightarrow$	4.00
Three SFs		Calculator display		Final answer, two zeros added to give three SFs

CONCEPT CHECK 1.6

Determination of Significant Figures for Final Answer

Perform the following calculation of measured numbers. Give the final answer with the correct number of significant figures.

$$\frac{8.560}{(2.84)(0.078)}$$

ANSWER

The calculation is performed as follows:

8.560	÷	2.84	÷	0.078	=	38.64210906	⟶	39
Four SFs		Three SFs		Two SFs		Calculator display		Final answer, rounded off to two SFs

The measurement with only two SFs limits the final answer to two SFs. To obtain the final answer, the calculator display must be rounded off by dropping all the digits after the 38. Because the first dropped digit 6 is greater than 5, the last retained digit is increased by one to give the final answer of 39.

SAMPLE PROBLEM 1.4

Significant Figures in Multiplication and Division

Perform the following calculations of measured numbers. Give the answers with the correct number of significant figures.

a. 56.8×0.37 **b.** $\dfrac{(2.075)(0.585)}{(8.42)(0.0045)}$ **c.** $\dfrac{25.0}{5.00}$

SOLUTION

a. 21 **b.** 32 **c.** 5.00 (add two significant zeros)

STUDY CHECK 1.4

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

a. $45.26 \times 0.010\,88$ **b.** $2.6 \div 324$ **c.** $\dfrac{4.0 \times 8.00}{16}$

Addition and Subtraction

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

Example 3

Add:

2.045	Three decimal places
+ 34.1	One decimal place
36.145	Calculator display
36.1	Answer, rounded off to one decimal place

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

Example 4

Subtract:

14.5 g	One decimal place
$\boxed{-}$ 2.5 g	One decimal place
12.	Calculator display
12.0 g	Answer, zero written after the decimal point

SAMPLE PROBLEM 1.5

Decimal Places in Addition and Subtraction

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

a. 27.8 cm + 0.235 cm **b.** 153.247 g − 14.82 g

SOLUTION

a. 28.0 cm (rounded off to one decimal place)
b. 138.43 g (rounded off to two decimal places)

STUDY CHECK 1.5

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

a. 82.45 mg + 1.245 mg + 0.000 56 mg **b.** 4.259 L − 3.8 L

QUESTIONS AND PROBLEMS

Significant Figures in Calculations

1.27 Why do we usually need to round off calculations that use measured numbers?

1.28 Why do we sometimes add a zero to a number in a calculator display?

1.29 Round off each of the following measurements to three significant figures:
a. 1.854 kg
b. 88.2038 L
c. 0.004 738 265 cm
d. 8807 m
e. 1.832×10^5 s

1.30 Round off each of the measurements in Problem 1.29 to two significant figures.

1.31 Round off or add zeros to each of the following to give an answer with three significant figures:
a. 5080 L **b.** 37 400 g
c. 104 720 m **d.** 0.000 250 82 s

1.32 Round off or add zeros to each of the following to give an answer with two significant figures:
a. 5 100 000 L **b.** 26 711 s
c. 0.003 378 m **d.** 56.982 g

1.33 For each of the following problems, give an answer with the correct number of significant figures:
a. 45.7×0.034 **b.** $0.002\ 78 \times 5$
c. $\dfrac{34.56}{1.25}$ **d.** $\dfrac{(0.2465)(25)}{1.78}$

1.34 For each of the following problems, give an answer with the correct number of significant figures:
a. 400×185 **b.** $\dfrac{2.40}{(4)(125)}$
c. $0.825 \times 3.6 \times 5.1$ **d.** $\dfrac{3.5 \times 0.261}{8.24 \times 20.0}$

1.35 For each of the following, give an answer with the correct number of decimal places:
a. 45.48 cm + 8.057 cm
b. 23.45 g + 104.1 g + 0.025 g
c. 145.675 mL − 24.2 mL
d. 1.08 L − 0.585 L

1.36 For each of the following, give an answer with the correct number of decimal places:
a. 5.08 g + 25.1 g
b. 85.66 cm + 104.10 cm + 0.025 cm
c. 24.568 mL − 14.25 mL
d. 0.2654 L − 0.2585 L

1.7 Prefixes and Equalities

LEARNING GOAL

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

SELF STUDY ACTIVITY
The Metric System

TUTORIAL
SI Prefixes and Units

The special feature of the SI as well as the metric system is that a **prefix** can be placed in front of any unit to increase or decrease its size by some factor of ten. For example, the prefixes *milli* and *micro* are used to make the smaller units, milligram (mg) and microgram (μg). Table 1.8 lists some of the metric prefixes, their symbols, and their numerical values.

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

The U.S. Food and Drug Administration has determined the daily values (DV) of nutrients for adults and children age four or older. Examples of these recommended daily values, some of which use prefixes, are listed in Table 1.9.

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

1 **kilo**meter (1 km) = **1000** meters (**10^3** m)

1 **kilo**liter (1 kL) = **1000** liters (**10^3** L)

1 **kilo**gram (1 kg) = **1000** grams (**10^3** g)

TABLE 1.8 Metric and SI Prefixes

Prefix	Symbol	Numerical Value	Scientific Notation	Equality
Prefixes That Increase the Size of the Unit				
tera	T	1 000 000 000 000	10^{12}	1 Ts = 1×10^{12} s 1 s = 1×10^{-12} Ts
giga	G	1 000 000 000	10^{9}	1 Gm = 1×10^{9} m 1 m = 1×10^{-9} Gm
mega	M	1 000 000	10^{6}	1 Mg = 1×10^{6} g 1 g = 1×10^{-6} Mg
kilo	k	1 000	10^{3}	1 km = 1×10^{3} m 1 m = 1×10^{-3} km
Prefixes That Decrease the Size of the Unit				
deci	d	0.1	10^{-1}	1 dL = 1×10^{-1} L 1 L = 10 dL
centi	c	0.01	10^{-2}	1 cm = 1×10^{-2} m 1 m = 100 cm
milli	m	0.001	10^{-3}	1 ms = 1×10^{-3} s 1 s = 1×10^{3} ms
micro	μ	0.000 001	10^{-6}	1 μg = 1×10^{-6} g 1 g = 1×10^{6} μg
nano	n	0.000 000 001	10^{-9}	1 nm = 1×10^{-9} m 1 m = 1×10^{9} nm
pico	p	0.000 000 000 001	10^{-12}	1 ps = 1×10^{-12} s 1 s = 1×10^{12} ps

TABLE 1.9 Daily Values for Selected Nutrients

Nutrient	Amount Recommended
Protein	44 g
Vitamin C	60 mg
Vitamin B_{12}	6 μg
Calcium	1000 mg
Copper	2 mg
Iodine	150 μg
Iron	18 mg
Magnesium	400 mg
Niacin	20 mg
Potassium	3500 mg
Selenium	55 μg
Sodium	2400 mg
Zinc	15 mg

CONCEPT CHECK 1.7

Prefixes

Fill in the blanks with the correct symbol.

a. 1000 g = 1 ___ g **b.** 0.01 m = 1 ___ m
c. 1×10^6 L = 1 ___ L

ANSWER

a. The prefix for 1000 is *kilo*; 1000 g = 1 kg
b. The prefix for 0.01 is *centi*; 0.01 m = 1 cm
c. The prefix for 1×10^6 is *mega*; 1×10^6 L = 1 ML

A terabyte hard disk drive stores 10^{12} bytes of information.

SAMPLE PROBLEM 1.6

Prefixes

The storage capacity for a hard disk drive (HDD) is specified using prefixes: megabyte (MB), gigabyte (GB), or terabyte (TB). Indicate the storage capacity in bytes for each of the following hard disk drives. Suggest a reason for describing a HDD storage capacity in gigabytes or terabytes.

a. 5 MB **b.** 2 GB

SOLUTION

a. The prefix *mega* (M) in MB is equal to 1 000 000 or 1×10^6. Thus, 5 MB is equal to 5 000 000 (5×10^6) bytes.
b. The prefix *giga* (G) in GB is equal to 1 000 000 000 or 1×10^9. Thus, 2 GB is equal to 2 000 000 000 (2×10^9) bytes.

Expressing HDD capacity in gigabytes or terabytes gives a more reasonable number to work with than a number with many zeros or a large power of 10.

STUDY CHECK 1.6

Hard drives now have a storage capacity of 1.5 TB. How many bytes are stored?

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

Measuring Length

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

If we compare the lengths of a millimeter and a centimeter, we find that 1 mm is 0.1 cm; there are 10 mm in 1 cm. These comparisons are examples of **equalities**, which show the relationship between two units that measure the same quantity. For example, in the equality 1 m = 100 cm, each quantity describes the same length but in a different unit. In every equality, each quantity has both a number and a unit.

First quantity			Second quantity	
1	m	=	100	cm
↑	↑		↑	↑
Number +	unit		Number +	unit

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

Examples of Some Length Equalities

1 m = 100 cm = 1×10^2 cm
1 m = 1000 mm = 1×10^3 mm
1 cm = 10 mm = 1×10^1 mm

Some metric units for length are compared in Figure 1.8.

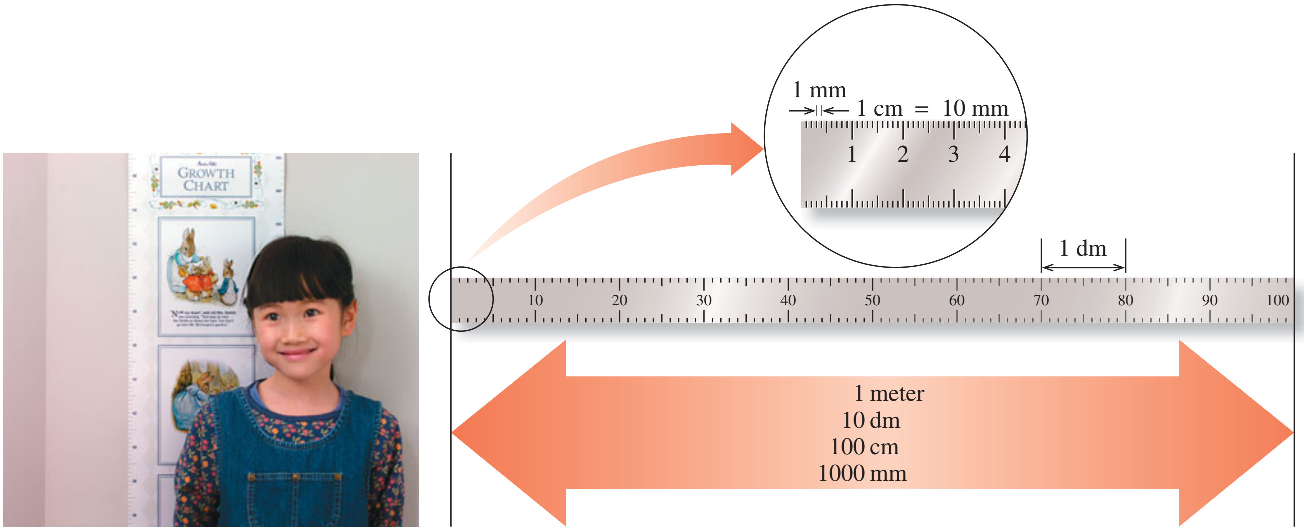


FIGURE 1.8 The metric length of 1 meter is the same length as 10 dm, 100 cm, or 1000 mm.

Q How many millimeters (mm) are in 1 centimeter (cm)?

Measuring Volume

Volumes of 1 L or smaller are common in the health sciences. When a liter is divided into 10 equal portions, each portion is a deciliter (dL). There are 10 dL in 1 L. Laboratory results for blood work are often reported in mass per deciliter. Table 1.10 lists typical laboratory test values for some substances in the blood.

When a liter is divided into a thousand parts, each of the smaller volumes is called a milliliter (mL). In a 1-L container of physiological saline, there are 1000 mL of solution (see Figure 1.9).

Examples of Some Volume Equalities

$1\ \text{L} = 10\ \text{dL} = 1 \times 10^1\ \text{dL}$

$1\ \text{L} = 1000\ \text{mL} = 1 \times 10^3\ \text{mL}$

$1\ \text{dL} = 100\ \text{mL} = 1 \times 10^2\ \text{mL}$

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

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

When you see *1 cm*, you are reading about length; when you see *1 cc* or *1 cm^3* or *1 mL*, you are reading about volume. A comparison of units of volume is illustrated in Figure 1.10.

Measuring Mass

When you go to the doctor for a physical examination, your mass is recorded in kilograms, whereas the results of your laboratory tests are reported in grams, milligrams (mg), or micrograms (μg). A kilogram is equal to 1000 g. One gram represents the same mass as 1000 mg, and one mg equals 1000 μg.

Examples of Some Mass Equalities

$1\ \text{kg} = 1000\ \text{g} = 1 \times 10^3\ \text{g}$

$1\ \text{g} = 1000\ \text{mg} = 1 \times 10^3\ \text{mg}$

$1\ \text{mg} = 1000\ \mu\text{g} = 1 \times 10^3\ \mu\text{g}$

TABLE 1.10 Some Typical Laboratory Test Values

Substance in Blood	Typical Range
Albumin	3.5–5.0 g/dL
Ammonia	20–150 μg/dL
Calcium	8.5–10.5 mg/dL
Cholesterol	105–250 mg/dL
Iron (male)	80–160 μg/dL
Protein (total)	6.0–8.0 g/dL

A laboratory technician transfers small volumes using a micropipette.

FIGURE 1.9 A plastic intravenous fluid container contains 1000 mL.

Q How many liters of solution are in the intravenous fluid container?

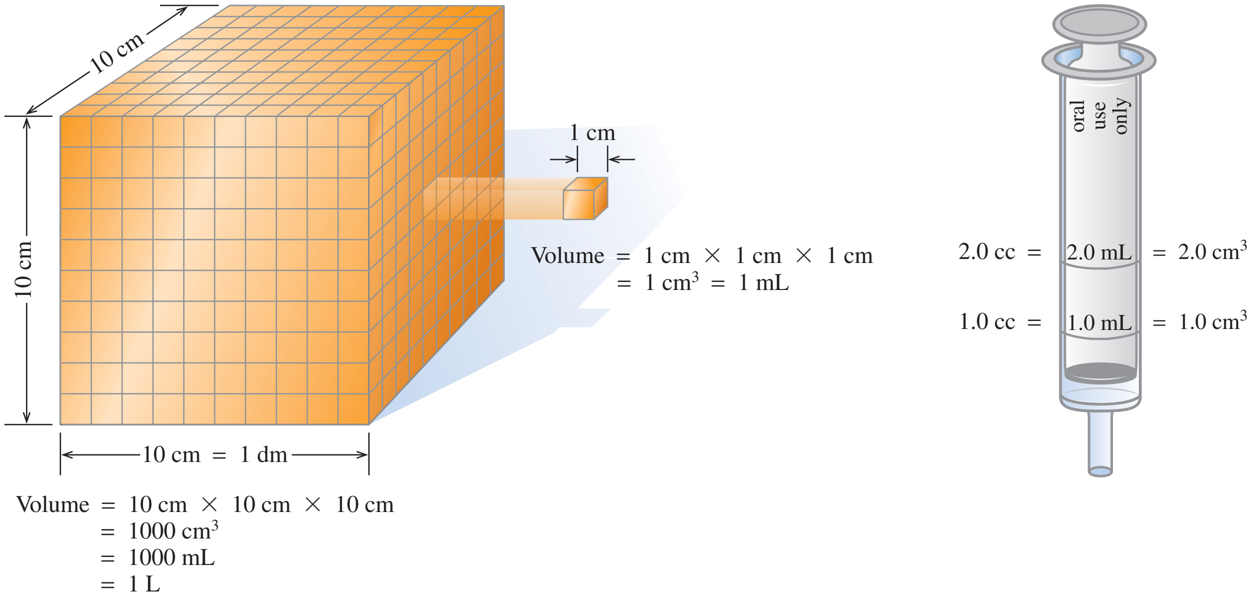


FIGURE 1.10 A cube measuring 10 cm on each side has a volume of 1000 cm^3, or 1 L; a cube measuring 1 cm on each side has a volume of 1 cm^3 (cc) or 1 mL.
Q What is the relationship between a milliliter (mL) and a cubic centimeter (cm^3)?

CONCEPT CHECK 1.8

Metric Relationships

Identify the larger unit in each of the following pairs:

a. centimeter or kilometer
b. mg or μg

ANSWER

a. A kilometer (1000 m) is larger than a centimeter (0.01 m).
b. A mg (0.001 g) is larger than a μg (0.000 001 g).

SAMPLE PROBLEM 1.7

Writing Metric Equalities

Complete the following list of metric equalities:

a. 1 L = ________ dL
b. 1 km = ________ m
c. 1 cm^3 = ________ mL

SOLUTION

a. 10 dL **b.** 1×10^3 m **c.** 1 mL

STUDY CHECK 1.7

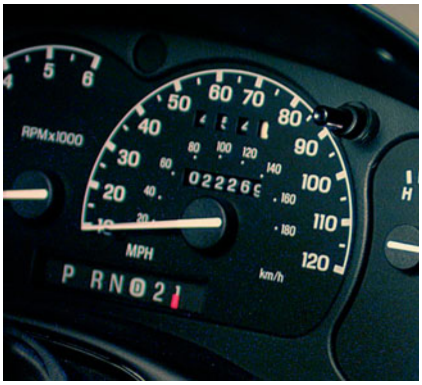
Complete each of the following equalities:

a. 1 kg = ________ g
b. 1 mL = ________ L

QUESTIONS AND PROBLEMS

Prefixes and Equalities

1.37 The speedometer is marked in both km/h and mi/h, or mph. What is the meaning of each abbreviation?



1.38 In a French car, the odometer reads 2250. What units would this be? What units would it be if this were an odometer in a car made for the United States?

1.39 How does the prefix *kilo* affect the gram unit in *kilogram*?

1.40 How does the prefix *centi* affect the meter unit in *centimeter*?

1.41 Write the abbreviation for each of the following units:
a. milligram **b.** deciliter
c. kilometer **d.** picogram
e. microliter **f.** nanosecond

1.42 Write the complete name for each of the following units:
a. cm **b.** ks
c. dL **d.** Gm
e. μg **f.** pg

1.43 Write the numerical values for each of the following prefixes:
a. centi **b.** tera
c. milli **d.** deci
e. mega **f.** pico

1.44 Write the complete name (prefix + unit) for each of the following numerical values:
a. 0.10 g **b.** 0.000 001 g
c. 1000 g **d.** 0.01 g
e. 0.001 g **f.** 1×10^9 g

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

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

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

1.48 For each of the following pairs, which is the smaller unit?
a. mg or g **b.** centimeter or millimeter
c. mm or μm **d.** mL or dL
e. mg or Mg

1.8 Writing Conversion Factors

LEARNING GOAL

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

TUTORIAL
Metric Conversions

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

Two Conversion Factors for the Equality 1 h = 60 min

$$\frac{\text{Numerator}}{\text{Denominator}} \begin{matrix}\longrightarrow\\ \longrightarrow\end{matrix} \quad \frac{60\text{ min}}{1\text{ h}} \quad \text{and} \quad \frac{1\text{ h}}{60\text{ min}}$$

These factors are read as "60 minutes per 1 hour" and "1 hour per 60 minutes." The term *per* means "divided by." Some common relationships are given in Table 1.11. It is important that the equality you select to form a conversion factor is a true relationship.

When an equality shows the relationship for two units from the same system (metric or U.S.), it is considered a definition and exact. Thus, the numbers in that definition are not used to determine significant figures. When an equality shows the relationship of units from two different systems, the number is measured and counts toward the significant figures in a calculation. For example, in the equality 1 lb = 454 g, the measured number

TABLE 1.11 Some Common Equalities

Quantity	U.S.	Metric (SI)	Metric–U.S.
Length	1 ft = 12 in. 1 yd = 3 ft 1 mi = 5280 ft	1 km = 1000 m 1 m = 1000 mm 1 cm = 10 mm	2.54 cm = 1 in. (exact) 1 m = 39.4 in. 1 km = 0.621 mi
Volume	1 qt = 4 cups 1 qt = 2 pt 1 gal = 4 qt	1 L = 1000 mL 1 dL = 100 mL 1 mL = 1 cm^3	946 mL = 1 qt 1 L = 1.06 qt
Mass	1 lb = 16 oz	1 kg = 1000 g 1 g = 1000 mg	1 kg = 2.20 lb 454 g = 1 lb
Time		1 h = 60 min 1 min = 60 s	

454 has three significant figures. The number one in 1 lb is considered as exact. An exception is the relationship of 1 in. = 2.54 cm: The value 2.54 has been defined as exact.

Metric Conversion Factors

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

Metric Equality	Conversion Factors
1 m = 100 cm	$\frac{100\text{ cm}}{1\text{ m}}$ and $\frac{1\text{ m}}{100\text{ cm}}$

Both are proper conversion factors for the relationship; one is just the inverse of the other. The usefulness of conversion factors is enhanced by the fact that we can turn a conversion factor over and use its inverse.

CONCEPT CHECK 1.9

Conversion Factors

Identify one or more correct conversion factors for the equality of gigagrams and grams.

a. $\frac{1\text{ Gg}}{1 \times 10^9\text{ g}}$ **b.** $\frac{1 \times 10^{-9}\text{ g}}{1\text{ Gg}}$ **c.** $\frac{1 \times 10^9\text{ Gg}}{1\text{ g}}$ **d.** $\frac{1 \times 10^9\text{ g}}{1\text{ Gg}}$

ANSWER

In Table 1.8, the equality for gigagrams and grams is 1 Gg = 1×10^9 g. Answers **a** and **d** are correctly written conversion factors.

FIGURE 1.11 In the United States, the contents of many packaged foods are listed in both U.S. and metric units.

Q What are some advantages of using the metric system?

Metric–U.S. System Conversion Factors

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

$$1\text{ kg} = 2.20\text{ lb}$$

The corresponding conversion factors would be

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

Figure 1.11 illustrates the contents of some packaged foods in both U.S. and metric units.

CONCEPT CHECK 1.10

Writing Conversion Factors from Equalities

Write conversion factors for the relationship between the following pairs of units:

a. millimeters and meters
b. quarts and milliliters

ANSWER

Equality	Conversion Factors
a. 1 m = 1000 mm	$\frac{1 \text{ m}}{1000 \text{ mm}}$ and $\frac{1000 \text{ mm}}{1 \text{ m}}$
b. 1 qt = 946 mL	$\frac{1 \text{ qt}}{946 \text{ mL}}$ and $\frac{946 \text{ mL}}{1 \text{ qt}}$

Conversion Factors Stated within a Problem

Many times, an equality is specified within a problem that applies only to that problem. It might be the price of 1 kilogram of oranges or the speed of a car in kilometers per hour. However, we can write conversion factors for relationships stated within a problem.

1. The motorcycle was traveling at a speed of 85 km/h.
   Equality 1 h = 85 km
   Conversion factors $\frac{85 \text{ km}}{1 \text{ h}}$ and $\frac{1 \text{ h}}{85 \text{ km}}$
2. One vitamin tablet contains 500 mg of vitamin C.
   Equality 1 tablet = 500 mg of vitamin C
   Conversion factors $\frac{500 \text{ mg vitamin C}}{1 \text{ tablet}}$ and $\frac{1 \text{ tablet}}{500 \text{ mg vitamin C}}$

Conversion Factors from Percent, ppm, and ppb

When a *percent* (%) is given in a problem, it means parts per 100 parts. To write a percent as a conversion factor, we choose a unit and express the numerical relationship of the parts of this unit to 100 parts of the whole. For example, a person might have 18% body fat by mass. The percent quantity can be written as 18 mass units of body fat in every 100 mass units of body mass. Different mass units such as grams (g), kilograms (kg), or pounds (lb) can be used, but both units in the factor must be the same.

Percent quantity 18% body fat by mass
Equality 18 kg of body fat = 100 kg of body mass
Conversion factors $\frac{100 \text{ kg body mass}}{18 \text{ kg body fat}}$ and $\frac{18 \text{ kg body fat}}{100 \text{ kg body mass}}$

When scientists want to indicate ratios with particularly small percentage values, they use numerical relationships called *parts per million* (ppm) or *parts per billion* (ppb). The ratio of parts per million is the same as the milligrams of a substance per kilogram (mg/kg). The ratio of parts per billion equals the micrograms per kilogram (μg/kg). For example, the maximum amount of lead allowed by the FDA in glazed pottery bowls is 5 ppm, which is 5 mg/kg.

ppm quantity 5 ppm of lead in glaze
Equality 5 mg of lead = 1 kg of glaze
Conversion factors $\frac{5 \text{ mg lead}}{1 \text{ kg glaze}}$ and $\frac{1 \text{ kg glaze}}{5 \text{ mg lead}}$

Explore Your World

SI AND METRIC EQUALITIES ON PRODUCT LABELS

Read the labels on some food products on your kitchen shelves and in your refrigerator, or refer to the labels in Figure 1.11. List the amount of product given in different units. Write a relationship for two of the amounts for the same product and container. Look for measurements of grams and pounds or quarts and milliliters.

QUESTIONS

1. Use the stated measurement to derive a metric–U.S. conversion factor.
2. How do your results compare with the conversion factors we have described in this text?

Vitamin C, an antioxidant needed by the body, is found in fruits such as lemons.

The thickness of the skin fold at the abdomen is used to determine percent body fat.

The maximum amount of mercury allowed in tuna is 0.1 ppm.

SAMPLE PROBLEM 1.8

Conversion Factors Stated in a Problem

Write two conversion factors for each of the following statements:

a. There are 325 mg of aspirin in 1 tablet.
b. The EPA has set the maximum level for mercury in tuna at 0.1 ppm.

SOLUTION

a. $\dfrac{325 \text{ mg aspirin}}{1 \text{ tablet}}$ and $\dfrac{1 \text{ tablet}}{325 \text{ mg aspirin}}$

b. $\dfrac{0.1 \text{ mg mercury}}{1 \text{ kg tuna}}$ and $\dfrac{1 \text{ kg tuna}}{0.1 \text{ mg mercury}}$

STUDY CHECK 1.8

What conversion factors can be written for the following statements?

a. A cyclist in the Tour de France bicycle race rides at the average speed of 62.2 km/h.
b. The permissible level of arsenic in water is 10 ppb.

Chemistry Link to Health

TOXICOLOGY AND RISK-BENEFIT ASSESSMENT

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

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

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

Dosage	Units
parts per million (ppm)	milligrams per kilogram (mg/kg)
parts per billion (ppb)	micrograms per kilogram (μg/kg)

Other evaluations need to be made, but it is easy to compare LD_{50} values. Parathion, a pesticide, with an LD_{50} of 3 mg/kg would be highly toxic. That means that 3 mg of parathion per kg of body mass would be fatal to half the test animals. Table salt (sodium chloride) with an LD_{50} of 3000 mg/kg would have a much lower toxicity. You would need to ingest a huge amount of salt before any toxic effect would be observed. Although the risk to animals can be evaluated in the laboratory, it is more difficult to determine the impact in the environment since there is also a difference between continued exposure and a single, large dose of the substance.

Table 1.12 lists some LD_{50} values and compares substances in order of increasing toxicity.

TABLE 1.12 Some LD_{50} Values for Substances Tested in Rats

Substance	LD_{50} (mg/kg)
Table sugar	29 700
Baking soda	4220
Table salt	3000
Ethanol	2080
Aspirin	1100
Caffeine	192
DDT	113
Sodium cyanide	6
Parathion	3

The LD_{50} of caffeine is 192 ppm.

QUESTIONS AND PROBLEMS

Writing Conversion Factors

1.49 Why can two conversion factors be written for the equality 1 m = 100 cm?

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

1.51 Write the equality and conversion factors for each of the following:
a. centimeters and meters
b. milligrams and grams
c. liters and milliliters
d. deciliters and milliliters
e. days in 1 week

1.52 Write the equality and conversion factors for each of the following:
a. centimeters and inches
b. pounds and kilograms
c. pounds and grams
d. quarts and liters
e. dimes in 1 dollar

1.53 Write the conversion factors for each of the following statements:
a. A bee flies at an average speed of 3.5 m per second.
b. The daily value for potassium is 3500 mg.
c. An automobile traveled 46.0 km on 1.0 gal of gasoline.
d. The label on a bottle reads 50 mg of Atenolol per tablet.
e. The pesticide level in plums was 29 ppb.
f. A low-dose aspirin contains 81 mg of aspirin per tablet.

1.54 Write the conversion factors for each of the following statements:
a. The label on a bottle reads 10 mg of furosemide per mL.
b. The daily value for iodine is 150 μg.
c. The nitrate level in well water was 32 ppm.
d. Gold jewelry contains 58% by mass gold.
e. The price of a gallon of gas is \$3.19.
f. One capsule of fish oil contains 360 mg of omega-3 fatty acids.

1.9 Problem Solving

LEARNING GOAL

Use conversion factors to change from one unit to another.

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

Given unit $\times$ one or more conversion factors $=$ needed unit

TUTORIAL
Unit Conversions

TUTORIAL
Metric Conversions

SAMPLE PROBLEM 1.9

Problem Solving Using Metric Factors

In radiological imaging such as PET or CT scans, dosages of pharmaceuticals are based on body mass. If a person weighs 164 lb, what is the body mass in kilograms?

SOLUTION

Step 1 State the given and needed quantities.

Given 164 lb **Need** kilograms

Step 2 Write a plan to convert the given unit to the needed unit. We see that the given unit is in the U.S. system of measurement and the needed unit in the metric system. Therefore, the conversion factor must relate the U.S. unit lb to the metric unit kg.

pounds → U.S.–Metric factor → kilograms

Step 3 State the equalities and conversion factors needed to cancel units.

$$1\text{ kg} = 2.20\text{ lb}$$

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

Guide to Problem Solving Using Conversion Factors

1 State the given and needed quantities.

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

3 State the equalities and conversion factors needed to cancel units.

4 Set up problem to cancel units and calculate answer.

Step 4 **Set up problem to cancel units and calculate answer.** Write the given, 164 lb, and set up the conversion factor with the unit lb in the denominator (bottom number) to cancel out the given unit (lb) in the numerator.

$$\underset{\text{Given}}{164\ \cancel{\text{lb}}} \times \underset{\text{Conversion factor}}{\boxed{\frac{\overset{\text{Unit for answer goes here}}{1\ \text{kg}}}{2.20\ \cancel{\text{lb}}}}} = \underset{\text{Answer}}{74.5\ \text{kg}}$$

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

$$\cancel{\text{lb}} \times \frac{\text{kg}}{\cancel{\text{lb}}} = \text{kg} \quad \text{Unit needed for answer}$$

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

$$\underset{\text{Three SFs}}{164} \times \underset{\text{Three SFs}}{\frac{1}{2.20}} = 164 \div 2.20 = \underset{\text{Calculator display}}{\boxed{74.54545455}} = \underset{\text{Three SFs (rounded off)}}{74.5}$$

The value of 74.5 combined with the unit, kg, gives the final answer of 74.5 kg. With few exceptions, answers to numerical problems contain a number and a unit.

STUDY CHECK 1.9

If 1890 mL of orange juice is prepared from orange juice concentrate, how many liters of orange juice is that?

Using Two or More Conversion Factors

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

CONCEPT CHECK 1.11

Canceling Units

Cancel the units in the following setup and give the unit of the final answer:

$$3.5\ \text{L} \times \frac{1 \times 10^3\ \text{mL}}{1\ \text{L}} \times \frac{0.48\ \text{g}}{1\ \text{mL}} \times \frac{1 \times 10^3\ \text{mg}}{1\ \text{g}} =$$

ANSWER

All matching units in the denominator and numerator cancel to give mg as the needed unit for the answer.

$$3.5\ \cancel{\text{L}} \times \frac{1 \times 10^3\ \cancel{\text{mL}}}{1\ \cancel{\text{L}}} \times \frac{0.48\ \cancel{\text{g}}}{1\ \cancel{\text{mL}}} \times \frac{1 \times 10^3\ \text{mg}}{1\ \cancel{\text{g}}} = \text{needed unit of mg}$$

SAMPLE PROBLEM 1.10

Problem Solving Using Two Factors

Synthroid is used as a replacement or supplemental therapy for diminished thyroid function. A doctor's order prescribed a dosage of 0.200 mg. If tablets in stock contain 50 μg of Synthroid, how many tablets are required to provide the prescribed medication?

SOLUTION

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

Given 0.200 mg of Synthroid **Need** number of tablets

Step 2 **Write a plan to convert the given unit to the needed unit.**

milligrams → Metric factor → micrograms → Clinical factor → number of tablets

Step 3 **State the equalities and conversion factors needed to cancel units.** In the problem, the information for the dosage is given as 50 μg per tablet. Using this as an equality along with the metric equality for milligrams and micrograms provides the following conversion factors:

$$1\text{ mg} = 1000\ \mu\text{g} \qquad \frac{1\text{ mg}}{1000\ \mu\text{g}} \text{ and } \frac{1000\ \mu\text{g}}{1\text{ mg}}$$

$$1\text{ tablet} = 50\ \mu\text{g} \qquad \frac{1\text{ tablet}}{50\ \mu\text{g}} \text{ and } \frac{50\ \mu\text{g}}{1\text{ tablet}}$$

Step 4 **Set up problem to cancel units and calculate answer.** The problem can be set up using the metric factor to cancel milligrams, and then the clinical factor to obtain tablets as the final unit.

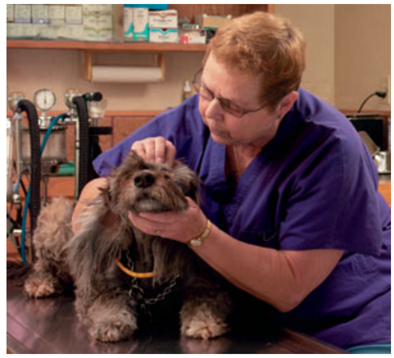
$$0.200\ \cancel{\text{mg}} \times \frac{1000\ \cancel{\mu\text{g}}}{1\ \cancel{\text{mg}}} \times \frac{1\text{ tablet}}{50\ \cancel{\mu\text{g}}} = 4\text{ tablets}$$

STUDY CHECK 1.10

One medium bran muffin contains 4.2 g of fiber. How many ounces (oz) of fiber are obtained by eating three medium bran muffins if 1 lb = 16 oz?
(*Hint:* number of muffins ⟶ g of fiber ⟶ lb ⟶ oz)

Career Focus

VETERINARY TECHNICIAN (VT)



"I am checking this dog's ears for foxtails and her eyes for signs of conjunctivitis," says Joyce Rhodes, veterinary assistant at the Sonoma Animal Hospital. "We always check a dog's teeth for tartar, because dental care is very important to the well-being of the animal. When I do need to give a medication to an animal, I use my chemistry to prepare the proper dose that the pet should take. Dosages may be in milligrams, grams, or milliliters."

As a member of the veterinary health care team, a veterinary technician (VT) assists a veterinarian in the care and handling of animals. A VT takes medical histories, collects specimens, performs laboratory procedures, prepares an animal for surgery, assists in surgical procedures, takes X-rays, talks with animal owners, and cleans teeth.

SAMPLE PROBLEM 1.11

Using a Percent as a Conversion Factor

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

SOLUTION

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

Given 155 lb of body weight; 16% body fat

Need kilograms of body fat

Step 2 **Write a plan to convert the given unit to the needed unit.**

lb of body weight → U.S.–Metric factor → kg of body mass → Percent factor → kg of body fat

Step 3 **State the equalities and conversion factors needed to cancel units.** One equality is the U.S.–metric factor for lb and kg. The second is the percent factor derived from the percentage information given in the problem.

TUTORIAL
Using Percentage
as a Conversion Factor

$$1 \text{ kg of body fat} = 2.20 \text{ lb of body fat} \qquad \frac{2.20 \text{ lb body fat}}{1 \text{ kg body fat}} \text{ and } \frac{1 \text{ kg body fat}}{2.20 \text{ lb body fat}}$$

$$16 \text{ kg of body fat} = 100 \text{ kg of body mass} \qquad \frac{16 \text{ kg body fat}}{100 \text{ kg body mass}} \text{ and } \frac{100 \text{ kg body mass}}{16 \text{ kg body fat}}$$

Step 4 **Set up problem to cancel units and calculate answer.** We can set up the problem using conversion factors to cancel each unit, starting with lb of body weight, until we obtain the final unit, kg of body fat, in the numerator. After we count the significant figures in the measured quantities, we write the needed answer with the proper number of significant figures.

$$155 \cancel{\text{lb body weight}} \times \frac{1 \cancel{\text{kg body mass}}}{2.20 \cancel{\text{lb body weight}}} \times \frac{16 \text{ kg body fat}}{100 \cancel{\text{kg body mass}}} = 11 \text{ kg of body fat}$$

Three SFs — Three SFs — Two SFs — Two SFs

STUDY CHECK 1.11

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

QUESTIONS AND PROBLEMS

Problem Solving

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

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

1.57 Use metric conversion factors to solve each of the following problems:
- **a.** The height of a student is 175 cm. How tall is the student in meters?
- **b.** A cooler has a volume of 5500 mL. What is the capacity of the cooler in liters?
- **c.** A Bee Hummingbird has a mass of 0.0018 kg. What is the mass of the hummingbird in grams?

1.58 Use metric conversion factors to solve each of the following problems:
- **a.** The daily value of phosphorus is 800 mg. How many grams of phosphorus are recommended?
- **b.** A glass of orange juice contains 0.85 dL of juice. How many milliliters of orange juice is that?
- **c.** A package of chocolate instant pudding contains 2840 mg of sodium. How many grams of sodium is that?

1.59 Solve each of the following problems using one or more conversion factors:
- **a.** A container holds 0.500 qt of liquid. How many milliliters of lemonade will it hold?
- **b.** What is the mass, in kilograms, of a person who weighs 145 lb?
- **c.** An athlete has 15% by mass body fat. What is the weight of fat, in pounds, of a 74-kg athlete?
- **d.** A plant fertilizer contains 15% by mass nitrogen (N). In a container of soluble plant food, there are 10.0 oz of fertilizer. How many grams of nitrogen are in the container?
- **e.** In a candy factory, the nutty chocolate bars contain 22.0% by mass pecans. If 5.0 kg of pecans were used for candy last Tuesday, how many pounds of nutty chocolate bars were made?

Agricultural fertilizers applied to a field provide nitrogen for plant growth.

1.60 Solve each of the following problems using one or more conversion factors:
- **a.** You need 4.0 ounces of a steroid ointment. If there are 16 oz in 1 lb, how many grams of ointment does the pharmacist need to prepare?
- **b.** During surgery, a patient receives 5.0 pints of plasma. How many milliliters of plasma were given? (1 quart = 2 pints)
- **c.** Wine is 12% (by volume) alcohol. How many milliliters of alcohol are in a 0.750 L bottle of wine?
- **d.** Blueberry high-fiber muffins contain 51% dietary fiber. If a package with a net weight of 12 oz contains six muffins, how many grams of fiber are in each muffin?
- **e.** A jar of crunchy peanut butter contains 1.43 kg of peanut butter. If you use 8.0% of the peanut butter for a sandwich, how many ounces of peanut butter did you take out of the container?

1.61 Using conversion factors, solve each of the following clinical problems:
- **a.** You have used 250 L of distilled water for a dialysis patient. How many gallons of water is that?
- **b.** A patient needs 0.024 g of a sulfa drug. There are 8-mg tablets in stock. How many tablets should be given?
- **c.** The daily dose of ampicillin for the treatment of an ear infection is 115 mg/kg of body weight. What is the daily dose for a 34-lb child?

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

a. The physician has ordered 1.0 g of tetracycline to be given every 6 hours to a patient. If your stock on hand is 500-mg tablets, how many will you need for 1 day's treatment?

b. An intramuscular medication is given at 5.00 mg/kg of body weight. If you give 425 mg of medication to a patient, what is the patient's weight in pounds?

c. A physician has ordered 0.50 mg of atropine, intramuscularly. If atropine were available as 0.10 mg/mL of solution, how many milliliters would you need to give?

1.10 Density

LEARNING GOAL

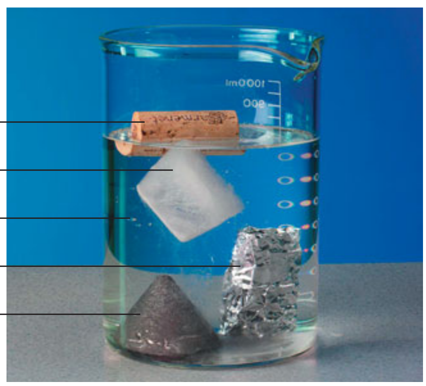
Calculate the density of a substance; use the density to calculate the mass or volume of a substance.

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

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

Every substance has a unique density, which distinguishes it from other substances. For example, lead has a density of 11.3 g/mL whereas cork has a density of 0.26 g/mL. From these densities, we can predict if these substances will sink or float in water. If a substance, such as cork, is less dense than water, it will float. However, a lead object sinks because its density is greater than that of water (see Figure 1.12).

Metals such as gold and lead tend to have higher densities whereas gases have low densities. In the metric system, the densities of solids and liquids are usually expressed as grams per cubic centimeter (g/cm^3) or grams per milliliter (g/mL). The densities of gases are usually stated as grams per liter (g/L). Table 1.13 gives the densities of some common substances.



Cork (D = 0.26 g/mL)
Ice (D = 0.92 g/mL)
Water (D = 1.00 g/mL)
Aluminum (D = 2.70 g/mL)
Lead (D = 11.3 g/mL)

FIGURE 1.12 Objects that sink in water are more dense than water; objects float if they are less dense.
Q Why does an ice cube float and a piece of aluminum sink?

TABLE 1.13 Densities of Some Common Substances

Solids (at 25 °C)	Density (g/mL)	Liquids (at 25 °C)	Density (g/mL)	Gases (at 0 °C)	Density (g/L)
Cork	0.26	Gasoline	0.74	Hydrogen	0.090
Wood (maple)	0.75	Ethanol	0.79	Helium	0.179
Ice (at 0 °C)	0.92	Olive oil	0.92	Methane	0.714
Sugar	1.59	Water (at 4 °C)	1.00	Neon	0.90
Bone	1.80	Urine	1.003–1.030	Nitrogen	1.25
Salt (NaCl)	2.16	Plasma (blood)	1.03	Air (dry)	1.29
Aluminum	2.70	Milk	1.04	Oxygen	1.43
Diamond	3.52	Mercury	13.6	Carbon dioxide	1.96
Copper	8.92				
Silver	10.5				
Lead	11.3				
Gold	19.3				

Explore Your World

SINK OR FLOAT?

1. Fill a large container or bucket with water. Place a can of diet soft drink and a can of a regular soft drink in the water. What happens? Using information on the label, how might you account for your observations?
2. Design an experiment to determine the substance that is the more dense in each of the following:
   a. water and vegetable oil
   b. water and ice
   c. rubbing alcohol and ice
   d. vegetable oil, water, and ice

CONCEPT CHECK 1.12

Density

(a) (b)

a. In drawing (a), the gray cube has a density of 4.5 g/cm^3. Is the density of the green cube the same, less than, or greater than that of the gray cube?

b. In drawing (b), the gray cube has a density of 4.5 g/cm^3. Is the density of the green cube the same, less than, or greater than that of the gray cube?

ANSWER

a. The green cube has the same volume as the gray cube but has a greater mass. Thus, the green cube has a density that is greater than that of the gray cube.

b. The green cube has the same mass as the gray cube, but the green cube has a greater volume. Thus, the green cube has a density that is less than that of the gray cube.

SAMPLE PROBLEM 1.12

Calculating Density

High-density lipoprotein (HDL) contains large amounts of proteins and small amounts of cholesterol. If a 0.258-g sample of HDL has a volume of 0.215 cm^3, what is the density of the HDL sample?

SOLUTION

Step 1 State the given and needed quantities.

Given mass of HDL sample = 0.258 g; volume = 0.215 cm^3

Need density (g/cm^3)

Step 2 Write the density expression.

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

Step 3 Express mass in grams and volume in milliliters (mL) or cm^3.

Mass of HDL sample = 0.258 g

Volume of HDL sample = 0.215 cm^3

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

$$\text{Density} = \frac{\overset{\text{Three SFs}}{0.258\text{ g}}}{\underset{\text{Three SFs}}{0.215\text{ cm}^3}} = \frac{1.20\text{ g}}{1\text{ cm}^3} = \underset{\text{Three SFs}}{1.20\text{ g/cm}^3}$$

STUDY CHECK 1.12

Low-density lipoprotein (LDL) contains small amounts of proteins and large amounts of cholesterol. If a 0.380-g sample of LDL has a volume of 0.362 cm^3, what is the density of the LDL sample?

Guide to Calculating Density

1 State the given and needed quantities.

2 Write the density expression.

3 Express mass in grams and volume in milliliters (mL) or cm^3.

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

Density Using Volume Displacement

The volume of a solid can be determined by volume displacement. When a solid is completely submerged in water, it displaces a volume that is equal to the volume of the solid. In Figure 1.13, the water level rises from 35.5 mL to 45.0 mL after the zinc object is added. This means that 9.5 mL of water is displaced and that the volume of the object is 9.5 mL. The density of the zinc is calculated as follows:

$$\text{Density} = \frac{\overset{\text{Four SFs}}{68.60\text{ g zinc}}}{\underset{\text{Two SFs}}{9.5\text{ mL}}} = \underset{\text{Two SFs}}{7.2\text{ g/mL}}$$

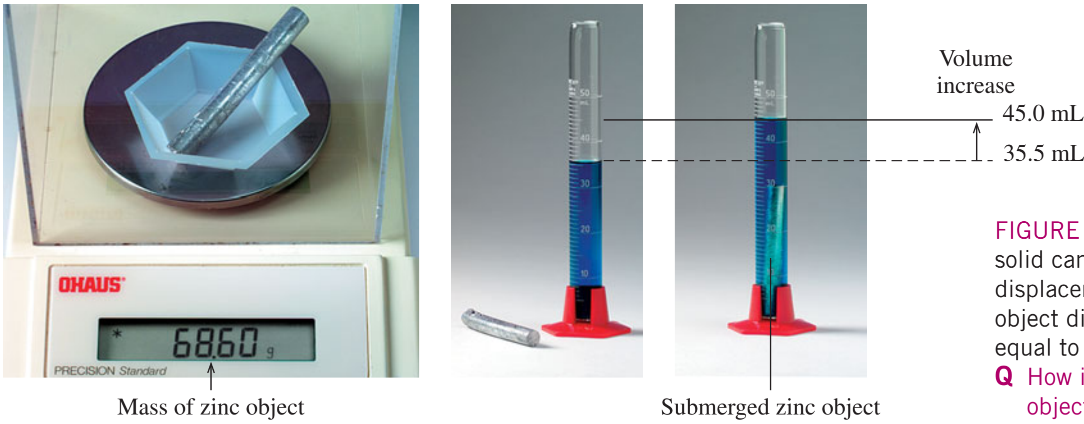


FIGURE 1.13 The density of a solid can be determined by volume displacement because a submerged object displaces a volume of water equal to its own volume.

Q How is the volume of the zinc object determined?

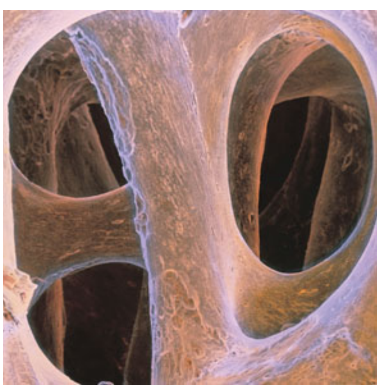
Chemistry Link to Health

BONE DENSITY

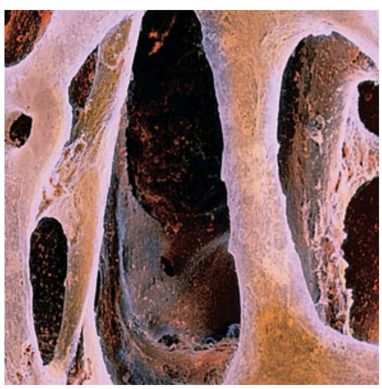
The density of our bones determines their health and strength. Our bones are constantly gaining and losing minerals such as calcium, magnesium, and phosphate. In childhood, bones form at a faster rate than they break down. As we age, the breakdown of bone occurs more rapidly than new bone forms. As the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

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

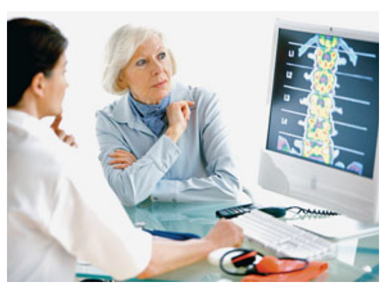
Recommendations to improve bone strength include supplements of calcium and vitamin D and medications such as Fosamax, Evista, or Actonel. Weight-bearing exercise such as walking and lifting weights can also improve muscle strength, which in turn, increases bone strength.



(**a**) Normal bone



(**b**) Bone with osteoporosis



(**c**) Viewing a low-dose X-ray of the spine

SAMPLE PROBLEM 1.13

Using Volume Displacement to Calculate Density

A lead weight used in the belt of a scuba diver has a mass of 226 g. When the lead weight is placed in a graduated cylinder containing 200.0 mL of water, the water level rises to 220.0 mL. What is the density of the lead weight (g/mL)?

Lead weights in a belt counteract the buoyancy of a scuba diver.

SOLUTION

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

Given mass = 226 g; water level before object submerged = 200.0 mL; water level after object submerged = 220.0 mL

Need density (g/mL)

Step 2 **Write the density expression.**

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

Step 3 **Express mass in grams and volume in milliliters (mL) or cm^3.**

Mass of lead weight = 226 g

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

$$\begin{array}{lcr} \text{Water level after object submerged} & = & 220.0\text{ mL} \\ \text{Water level before object submerged} & = & -200.0\text{ mL} \\ \hline \text{Water displaced (volume of lead weight)} & = & 20.0\text{ mL} \end{array}$$

Step 4 **Substitute mass and volume into the density expression and calculate the density.** The density is calculated by dividing the mass (g) by the volume (mL). Be sure to use the volume of water the object displaced and *not* the original volume of water.

$$\text{Density} = \frac{\overset{\text{Three SFs}}{226\text{ g}}}{\underset{\text{Three SFs}}{20.0\text{ mL}}} = \frac{11.3\text{ g}}{1\text{ mL}} = \underset{\text{Three SFs}}{11.3\text{ g/mL}}$$

STUDY CHECK 1.13

A total of 0.50 lb of glass marbles is added to 425 mL of water. The water level rises to a volume of 528 mL. What is the density (g/mL) of the glass marbles?

Guide to Using Density

1. State the given and needed quantities.
2. Write a plan to calculate the needed quantity.
3. Write equalities and their conversion factors including density.
4. Set up problem to calculate the needed quantity.

Problem Solving Using Density

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

SAMPLE PROBLEM 1.14

Problem Solving with Density

John took 2.0 teaspoons (tsp) of cough syrup for a persistent cough. If the syrup had a density of 1.20 g/mL and there is 5.0 mL in 1 tsp, what was the mass, in grams, of the cough syrup?

SOLUTION

Step 1 State the given and needed quantities.

Given 2.0 tsp **Need** grams of syrup

Step 2 Write a plan to calculate the needed quantity.

teaspoons → U.S.–Metric factor → milliliters → Density factor → grams

Step 3 Write equalities and their conversion factors including density.

$$1\text{ tsp} = 5.0\text{ mL} \qquad \frac{5.0\text{ mL}}{1\text{ tsp}} \text{ and } \frac{1\text{ tsp}}{5.0\text{ mL}}$$

$$1\text{ mL} = 1.20\text{ g of syrup} \qquad \frac{1\text{ mL}}{1.20\text{ g syrup}} \text{ and } \frac{1.20\text{ g syrup}}{1\text{ mL}}$$

Step 4 Set up problem to calculate the needed quantity.

$$2.0\ \cancel{\text{tsp}} \times \frac{5.0\ \cancel{\text{mL}}}{1\ \cancel{\text{tsp}}} \times \frac{1.20\text{ g syrup}}{1\ \cancel{\text{mL}}} = 12\text{ g of syrup}$$

STUDY CHECK 1.14

How many milliliters of mercury are in a thermometer that contains 20.4 g of mercury? (See Table 1.13 for the density of mercury.)

FIGURE 1.14 When the specific gravity of beer measures 1.010 or less with a hydrometer, the fermentation process is complete.
Q If the hydrometer reading is 1.006, what is the density of the liquid?

Specific Gravity

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

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

Specific gravity is one of the few unitless values you will encounter in chemistry. An instrument called a hydrometer is often used to measure the specific gravity of fluids such as battery fluid or a sample of urine. In Figure 1.14, a hydrometer is used to measure the specific gravity of beer.

QUESTIONS AND PROBLEMS

Density

1.63 In an old trunk, you find a piece of metal that you think may be aluminum, silver, or lead. In lab, you find it has a mass of 217 g and a volume of 19.2 cm^3. Using Table 1.13, what is the metal you found?

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

1.65 What is the density (g/mL) of each of the following samples?
a. A 20.0-mL sample of a salt solution that has a mass of 24.0 g.
b. A solid object with a mass of 1.65 lb and a volume of 170 mL.
c. A gem has a mass of 45.0 g. When the gem is placed in a graduated cylinder containing 20.0 mL of water, the water level rises to 34.5 mL.
d. A lightweight head on the driver of a golf club is made of titanium. If the volume of a sample of titanium is 114 cm^3 and the mass is 514.1 g, what is the density of titanium?

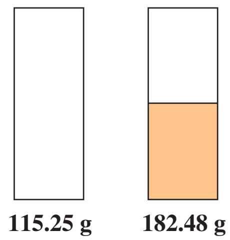
1.66 What is the density (g/mL) of each of the following samples?

a. A medication, if 3.00 mL has a mass of 3.85 g.

b. The fluid in a car battery, if it has a volume of 125 mL and a mass of 155 g.

c. A 5.00-mL urine sample from a patient suffering from symptoms resembling those of diabetes mellitus. The mass of the urine sample is 5.025 g.

d. A syrup is added to an empty container with a mass of 115.25 g. When 0.100 pint of syrup is added, the total mass of the container and syrup is 182.48 g. (1 qt = 2 pt)



1.67 Use the density value to solve the following problems:

a. What is the mass, in grams, of 150 mL of a liquid with a density of 1.4 g/mL?

b. What is the mass of a glucose solution that fills a 0.500-L intravenous bottle if the density of the glucose solution is 1.15 g/mL?

c. A sculptor has prepared a mold for casting a bronze figure. The figure has a volume of 225 mL. If bronze has a density of 7.8 g/mL, how many ounces of bronze are needed in the preparation of the bronze figure?

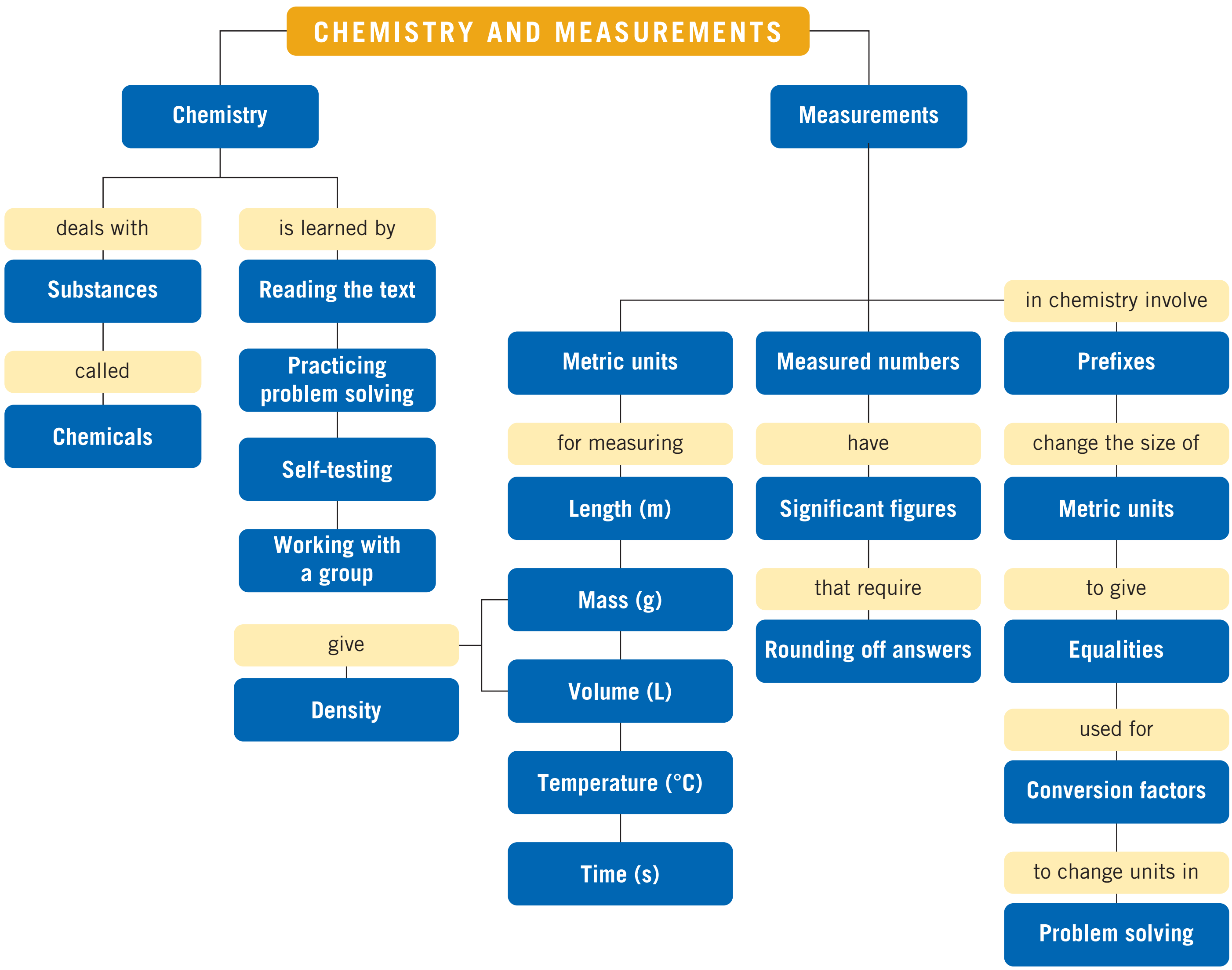
1.68 Use the density value to solve the following problems:

a. A graduated cylinder contains 18.0 mL of water. What is the new water level after 35.6 g of silver metal with a density of 10.5 g/mL is submerged in the water?

b. A thermometer containing 8.3 g of mercury has broken. If mercury has a density of 13.6 g/mL, what volume spilled?

c. A fish tank holds 35 gal of water. Using the density of 1.00 g/mL for water, determine the number of pounds of water in the fish tank.

CONCEPT MAP



CHAPTER REVIEW

1.1 Chemistry and Chemicals

Learning Goal: Define the term *chemistry* and identify substances as chemicals.

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

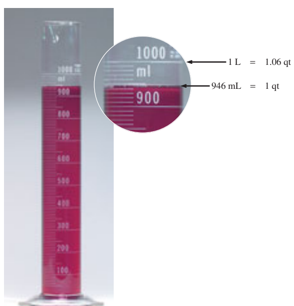
1.2 Study Plan for Learning Chemistry

Learning Goal: Develop a study plan for learning chemistry.

A study plan for learning chemistry utilizes the features in this text and develops an active learning approach to study chemistry. By using the *Learning Goals* in the chapter and working the *Concept Checks*, *Sample Problems*, *Study Checks*, and the *Questions and Problems* at the end of each section, the student can successfully learn the concepts of chemistry.

1.3 Units of Measurement

Learning Goal: Write the names and abbreviations for metric or SI units used in measurements of length, volume, mass, temperature, and time.

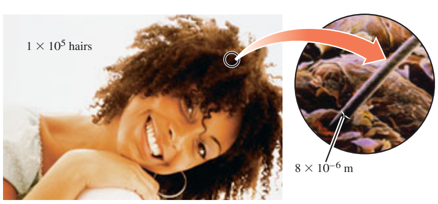


In science, physical quantities are described in units of the metric or International System (SI). Some important units are meter (m) for length, liter (L) for volume, gram (g) and kilogram (kg) for mass, degree Celsius (°C) and kelvin (K) for temperature, and second (s) for time.

1.4 Scientific Notation

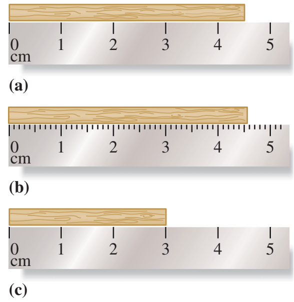
Learning Goal: Write a number in scientific notation.

Large and small numbers can be written using scientific notation, in which the decimal point is moved to give a coefficient of at least 1 but less than 10 and the number of decimal places moved shown as a power of 10. A large number will have a positive power of 10, while a small number will have a negative power of 10.



1.5 Measured Numbers and Significant Figures

Learning Goal: Identify a number as measured or exact; determine the number of significant figures in a measured number.



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

1.6 Significant Figures in Calculations

Learning Goal: Adjust calculated answers to give the correct number of significant figures.

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

1.7 Prefixes and Equalities

Learning Goal: Use the numerical values of prefixes to write a metric equality.

Prefixes placed in front of a unit change the size of the unit by factors of 10. Prefixes such as *centi*, *milli*, and *micro* provide smaller units; prefixes such as *kilo* provide larger units. An equality relates two units that measure the same quantity of length, volume, mass, or time. Examples of equalities are 1 m = 100 cm, 1 qt = 946 mL; 1 kg = 1000 g; 1 min = 60 s.

Examples of Some Mass Equalities

$1\ \text{kg} = 1000\ \text{g} = 1 \times 10^3\ \text{g}$

$1\ \text{g} = 1000\ \text{mg} = 1 \times 10^3\ \text{mg}$

$1\ \text{mg} = 1000\ \mu\text{g} = 1 \times 10^3\ \mu\text{g}$

1.8 Writing Conversion Factors

Learning Goal: Write a conversion factor for two units that describe the same quantity.

Conversion factors are used to express a relationship in the form of a fraction. Two factors can be written for any relationship in the metric or U.S. system.

1.9 Problem Solving

Learning Goal: Use conversion factors to change from one unit to another.

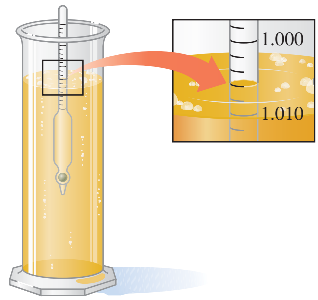
Conversion factors are useful when changing a quantity expressed in one unit to a quantity expressed in another unit. In the process, a given unit is multiplied by one or more conversion factors that cancel units until the needed answer is obtained.

Unit for answer goes here

$$\underbrace{164\ \cancel{\text{lb}}}_{\text{Given}} \times \underbrace{\boxed{\frac{1\ \text{kg}}{2.20\ \cancel{\text{lb}}}}}_{\text{Conversion factor}} = \underbrace{74.5\ \text{kg}}_{\text{Answer}}$$

1.10 Density

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



The density of a substance is a ratio of its mass to its volume, usually in units of g/mL or g/cm^3. Density can be used as a factor to convert between the mass and volume of a substance. Specific gravity (sp gr) compares the density of a substance to the density of water, 1.00 g/mL.

Key Terms

Celsius (°C) temperature scale A temperature scale on which water has a freezing point of 0 °C and a boiling point of 100 °C.

centimeter (cm) A unit of length in the metric system; there are 2.54 cm in 1 in.

chemical A substance that has the same composition and properties wherever it is found.

chemistry The science that studies the composition, structure, properties, and reactions of matter.

conversion factor A ratio in which the numerator and denominator are quantities from an equality or given relationship. For example, the conversion factors for the relationship 1 kg = 2.20 lb are written as

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

cubic centimeter (cm^3, cc) The volume of a cube that has 1-cm sides; 1 cm^3 is equal to 1 mL.

density The relationship of the mass of an object to its volume expressed as grams per cubic centimeter (g/cm^3), grams per milliliter (g/mL), or grams per liter (g/L).

equality A relationship between two units that measure the same quantity.

exact number A number obtained by counting or by definition.

gram (g) The metric unit used in measurements of mass.

International System of Units (SI) An international system of units that modifies the metric system.

Kelvin (K) temperature scale A temperature scale on which the lowest possible temperature is 0 K.

kilogram (kg) A metric mass of 1000 g, equal to 2.20 lb. The kilogram is the SI standard unit of mass.

liter (L) The metric unit for volume that is slightly larger than a quart.

mass A measure of the quantity of material in an object.

measured number A number obtained when a quantity is determined by using a measuring device.

meter (m) The metric unit for length that is slightly longer than a yard. The meter is the SI standard unit of length.

metric system A system of measurement used by scientists and in most countries of the world.

milliliter (mL) A metric unit of volume equal to one-thousandth of a liter (0.001 L).

prefix The part of the name of a metric unit that precedes the base unit and specifies the size of the measurement. All prefixes are related on a decimal scale.

scientific notation A form of writing large and small numbers using a coefficient that is at least 1 but less than 10, followed by a power of 10.

second (s) A unit of time used in both the SI and metric systems.

significant figures (SFs) The numbers recorded in a measurement.

temperature An indicator of the hotness or coldness of an object.

volume The amount of space occupied by a substance.

Understanding the Concepts

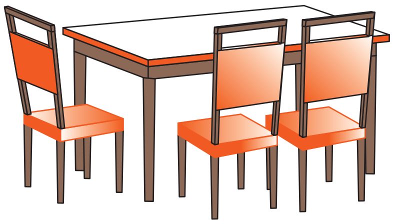
1.69 Which of the following will help you develop a successful study plan?
a. Skip lecture and just read the text.
b. Work the *Sample Problems* as you go through a chapter.
c. Go to your professor's office hours.
d. Read through the chapter, but work the problems later.

1.70 Which of the following will help you develop a successful study plan?
a. Study all night before the exam.
b. Form a study group and discuss the problems together.
c. Work problems in a notebook for easy reference.
d. Copy the answers to homework from a friend.

1.71 In which of the following pairs do both numbers contain the same number of significant figures?
a. 2.0500 m and 0.0205 m
b. 600.0 K and 60 K
c. 0.000 75 s and 75 000 s
d. 6.240 L and 6.240×10^{-2} L

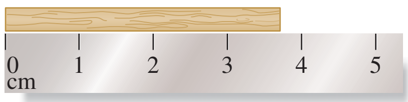
1.72 In which of the following pairs do both numbers contain the same number of significant figures?
a. 3.44×10^{-3} g and 0.0344 g
b. 0.0098 s and 9.8×10^{4} s
c. 6.8×10^{3} mm and 68 000 m
d. 258.000 ng and 2.58×10^{-2} g

1.73 Indicate if each of the following is answered with an exact number or a measured number:

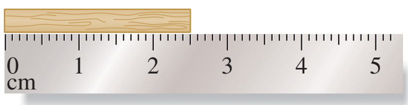


a. number of legs
b. height of table
c. number of chairs at the table
d. area of tabletop

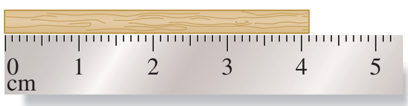
1.74 Measure the length of each of the objects in diagrams **(a)**, **(b)**, and **(c)** using the metric ruler in the figure. Indicate the number of significant figures for each and the estimated digit for each.



(a)

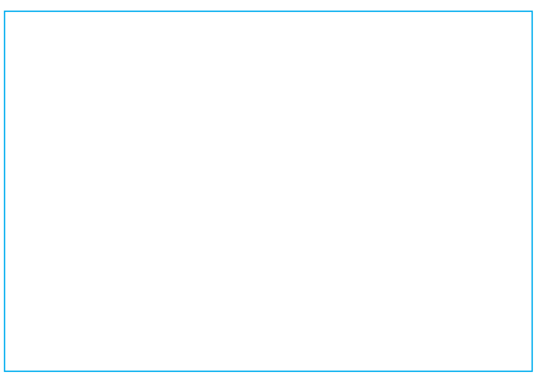


(b)



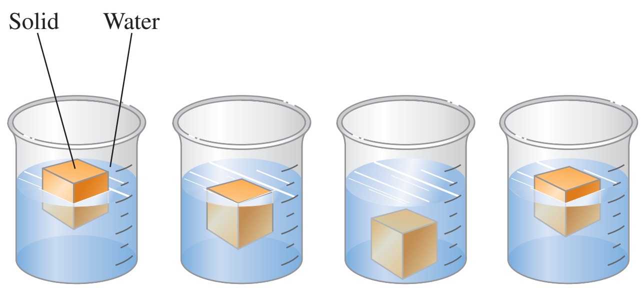
(c)

1.75 Measure the length and width of the rectangle, including the estimated digit, using a metric ruler.



a. What are the length and width of this rectangle measured in centimeters?
b. What are the length and width of this rectangle measured in millimeters?
c. How many significant figures are in the length measurement?
d. How many significant figures are in the width measurement?
e. What is the area of the rectangle in cm^2?
f. How many significant figures are in the calculated answer for area?

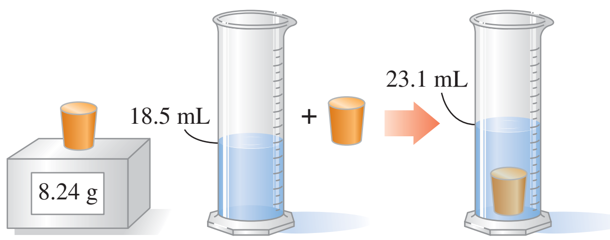
1.76 Each of the following diagrams represents a container of water and a cube. Some cubes float while others sink. Match diagrams 1, 2, 3, or 4 with one of the following descriptions and explain your choices:



1 **2** **3** **4**

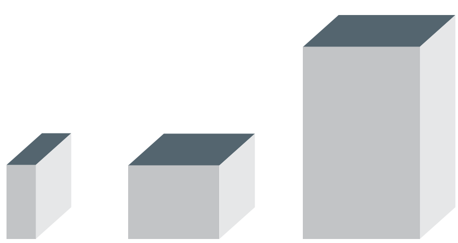
a. The cube has a greater density than water.
b. The cube has a density that is 0.60–0.80 g/mL.
c. The cube has a density that is one-half the density of water.
d. The cube has the same density as water.

1.77 What is the density of the solid object that is weighed and submerged in water?



1.78 Consider the following solids. The solids A, B, and C represent aluminum, gold, and silver. If each has a mass of 10.0 g, what is the identity of each solid?

Density of aluminum = 2.70 g/mL
Density of gold = 19.3 g/mL
Density of silver = 10.5 g/mL



A **B** **C**

Additional Questions and Problems

For instructor-assigned homework, go to **www.masteringchemistry.com**.

1.79 Round off or add zeros to the following calculated answers to give a final answer with three significant figures:
a. 0.000 012 58 L **b.** 3.528×10^2 kg
c. 125 111 m **d.** 58.703 g
e. 3×10^{-3} s **f.** 0.010 826 g

1.80 What is the total mass, in grams, of a dessert containing 137.25 g of vanilla ice cream, 84 g of fudge sauce, and 43.7 g of nuts?

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

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

1.83 In France, grapes are 1.75 Euros per kilogram. What is the cost of grapes, in dollars per pound, if the exchange rate is 1.36 dollars/Euro?

1.84 In Mexico, avocados are 48 pesos per kilogram. What is the cost, in cents, of an avocado that weighs 0.45 lb if the exchange rate is 13 pesos to the dollar?

1.85 Bill's recipe for onion soup calls for 4.0 lb of thinly sliced onions. If an onion has an average mass of 115 g, how many onions does Bill need?

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

1.87 The following nutrition information is listed on a box of crackers:

Serving size 0.50 oz (6 crackers)

Fat 4 g per serving Sodium 140 mg per serving

a. If the box has a net weight (contents only) of 8.0 oz, about how many crackers are in the box?
b. If you ate 10 crackers, how many ounces of fat are you consuming?
c. How many grams of sodium are used to prepare 50 boxes of crackers in part **a**?

1.88 A dialysis unit requires 75 000 mL of distilled water. How many gallons of water are needed? (1 gal = 4 qt)

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

1.90 Celeste's diet restricts her intake of protein to 24 g per day. If she eats 1.2 oz of protein, has she exceeded her protein limit for the day?

1.91 What is a cholesterol level of 1.85 g/L in the standard units of mg/dL?

1.92 An object has a mass of 3.15 oz and a volume of 0.1173 L. What is the density (g/mL) of the object?

1.93 The density of lead is 11.3 g/mL. The water level in a graduated cylinder initially at 215 mL rises to 285 mL after a piece of lead is submerged. What is the mass, in grams, of the lead?

1.94 A graduated cylinder contains 155 mL of water. A 15.0-g piece of iron (density = 7.86 g/cm^3) and a 20.0-g piece of lead (density = 11.3 g/cm^3) are added. What is the new water level, in milliliters, in the cylinder?

1.95 How many cubic centimeters (cm^3) of olive oil have the same mass as 1.50 L of gasoline (see Table 1.13)?

1.96 Ethyl alcohol has a density of 0.79 g/mL. What is the volume, in quarts, of 1.50 kg of alcohol?

Challenge Questions

1.97 A balance measures mass to 0.001 g. If you determine the mass of an object that weighs about 30 g, would you record the mass as 30 g, 32.5 g, 31.25 g, 31.075 g, or 3000 g? Explain your choice by writing two to three complete sentences that describe your thinking.

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

1.99 A car travels at 55 miles per hour and gets 11 kilometers per liter of gasoline. How many gallons of gasoline are needed for a 3.0-hour trip?

1.100 A 50.0-g silver object and a 50.0-g gold object are both added to 75.5 mL of water contained in a graduated cylinder. What is the new water level in the cylinder?

1.101 In the manufacturing of computer chips, cylinders of silicon are cut into thin wafers that are 3.00 inches in diameter and have a mass of 1.50 g of silicon. How thick (mm) is each wafer if silicon has a density of 2.33 g/cm^3? (The volume of a cylinder is $V = \pi r^2 h$.)

1.102 A sunscreen preparation contains 2.50% by mass benzyl salicylate. If a tube contains 4.0 ounces of sunscreen, how many kilograms of benzyl salicylate are needed to manufacture 325 tubes of sunscreen?

1.103 For a 180-lb person, calculate the quantities of each of the following that must be ingested to provide the LD_{50} for caffeine given in Table 1.12:
a. 100. mg caffeine per 6.0 fl oz
b. 50. mg of caffeine
c. 100. mg of caffeine

1.104 The label on a 1-pint bottle of mineral water lists the following components. If the density is the same as pure water and you drink 3 bottles of water in one day, how many milligrams of each component will you obtain?
a. calcium 28 ppm **b.** fluoride 0.08 ppm
c. magnesium 12 ppm **d.** potassium 3.2 ppm
e. sodium 15 ppm

1.105 **a.** Some athletes have as little as 3.0% body fat. If such a person has a body mass of 65 kg, how many pounds of body fat does that person have?

b. In liposuction, a doctor removes fat deposits from a person's body. If body fat has a density of 0.94 g/mL and 3.0 L of fat is removed, how many pounds of fat were removed from the patient?

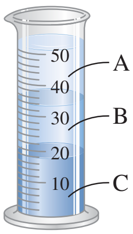
1.106 An 18-karat gold necklace is 75% gold by mass, 16% silver, and 9.0% copper.

a. What is the mass, in grams, of the necklace if it contains 0.24 oz of silver?

b. How many grams of copper are in the necklace?

c. If 18-karat gold has a density of 15.5 g/cm^3, what is the volume in cubic centimeters?

1.107 A graduated cylinder contains three liquids A, B, and C, which have different densities and do not mix: mercury (D = 13.6 g/mL), vegetable oil (D = 0.92 g/mL), and water (D = 1.00 g/mL). Identify the liquids A, B, and C in the cylinder.



1.108 A mouthwash is 21.6% ethyl alcohol by mass. If each bottle contains 0.358 pt of mouthwash with a density of 0.876 g/mL, how many kilograms of ethyl alcohol are in 180 bottles of the mouthwash?

A mouthwash may contain ethyl alcohol.

Answers

Answers to Study Checks

1.1 **a.** iron, **b.** tin, and **d.** water are chemicals.

1.2 **a.** 4.25×10^5 m **b.** 8.6×10^{-7} g

1.3 **a.** 36 m **b.** 0.0026 L **c.** 3.8×10^3 g **d.** 1.3 kg

1.4 **a.** 0.4924 **b.** 0.0080 or 8.0×10^{-3} **c.** 2.0

1.5 **a.** 83.70 mg **b.** 0.5 L

1.6 1 500 000 000 000 (1.5×10^{12}) bytes

1.7 **a.** 1000 **b.** 0.001

1.8 **a.** $\frac{62.2\text{ km}}{1\text{ h}}$ and $\frac{1\text{ h}}{62.2\text{ km}}$ **b.** $\frac{10\ \mu g}{1\text{ kg}}$ and $\frac{1\text{ kg}}{10\ \mu g}$

1.9 1.89 L

1.10 0.44 oz

1.11 25 g of fat

1.12 1.05 g/cm^3

1.13 2.2 g/mL

1.14 1.50 mL

Answers to Selected Questions and Problems

1.1 Many chemicals are listed on a vitamin bottle such as vitamin A, vitamin B_3, vitamin B_{12}, folic acid, and so on.

1.3 No. All of the ingredients listed are chemicals.

1.5 Among the things you might do to help yourself succeed in chemistry: attend lecture regularly, review the *Learning Goals*, keep a problem notebook, read the text actively, read the chapter before lecture, join a study group, use your instructor's office hours, and others.

1.7 **a**, **c**, **e**, and **f**

1.9 In the United States, **a.** body mass is measured in pounds, **b.** height in feet and inches, **c.** amount of gasoline in gallons, and **d.** temperature in degrees Fahrenheit (°F). In Mexico, **a.** body mass is measured in kilograms, **b.** height in meters, **c.** amount of gasoline in liters, and **d.** temperature in degrees Celsius (°C).

1.11 **a.** meter, length **b.** gram, mass **c.** milliliter, volume **d.** second, time **e.** degree Celsius, temperature

1.13 **a.** 5.5×10^4 m **b.** 4.8×10^2 g **c.** 5×10^{-6} cm **d.** 1.4×10^{-4} s **e.** 7.2×10^{-3} L **f.** 6.7×10^5 kg

1.15 **a.** 7.2×10^3 cm **b.** 3.2×10^{-2} kg **c.** 1×10^4 L **d.** 6.8×10^{-2} m

1.17 **a.** measured **b.** exact **c.** exact **d.** measured

1.19 **a.** 6 oz of hamburger **b.** none **c.** 0.75 lb, 350 g **d.** none (definitions are exact)

1.21 **a.** not significant **b.** significant **c.** significant **d.** significant **e.** not significant

1.23 **a.** 5 SFs **b.** 2 SFs **c.** 2 SFs **d.** 3 SFs **e.** 4 SFs **f.** 3 SFs

1.25 **c**

1.27 A calculator often gives more digits than the number of significant figures allowed in the answer.

1.29 **a.** 1.85 kg **b.** 88.2 L **c.** 0.004 74 cm **d.** 8810 m **e.** 1.83×10^5 s

1.31 **a.** 5.08×10^3 L **b.** 3.74×10^4 g **c.** 1.05×10^5 m **d.** 2.51×10^{-4} s

1.33 **a.** 1.6 **b.** 0.01 **c.** 27.6 **d.** 3.5

1.35 **a.** 53.54 cm **b.** 127.6 g **c.** 121.5 mL **d.** 0.50 L

1.37 km/h is kilometers per hour; mi/h is miles per hour

1.39 The prefix *kilo* means to multiply by 1000. One kg is the same mass as 1000 g.

1.41 **a.** mg **b.** dL **c.** km **d.** pg **e.** μL **f.** ns

1.43 **a.** 0.01 **b.** 1×10^{12} **c.** 1×10^{-3} **d.** 0.1 **e.** 1×10^6 **f.** 1×10^{-12}

1.45 **a.** 100 cm **b.** 1×10^9 nm **c.** 0.001 m **d.** 1000 mL

1.47 **a.** kilogram **b.** milliliter **c.** km **d.** kL **e.** nanometer

1.49 A conversion factor can be inverted to give a second conversion factor.

1.51 **a.** 100 cm = 1 m, $\dfrac{100\text{ cm}}{1\text{ m}}$ and $\dfrac{1\text{ m}}{100\text{ cm}}$

b. 1000 mg = 1 g, $\dfrac{1000\text{ mg}}{1\text{ g}}$ and $\dfrac{1\text{ g}}{1000\text{ mg}}$

c. 1 L = 1000 mL, $\dfrac{1000\text{ mL}}{1\text{ L}}$ and $\dfrac{1\text{ L}}{1000\text{ mL}}$

d. 1 dL = 100 mL, $\dfrac{100\text{ mL}}{1\text{ dL}}$ and $\dfrac{1\text{ dL}}{100\text{ mL}}$

e. 1 week = 7 days, $\dfrac{1\text{ week}}{7\text{ days}}$ and $\dfrac{7\text{ days}}{1\text{ week}}$

1.53 **a.** $\dfrac{3.5\text{ m}}{1\text{ s}}$ and $\dfrac{1\text{ s}}{3.5\text{ m}}$

b. $\dfrac{3500\text{ mg potassium}}{1\text{ day}}$ and $\dfrac{1\text{ day}}{3500\text{ mg potassium}}$

c. $\dfrac{46.0\text{ km}}{1.0\text{ gal}}$ and $\dfrac{1.0\text{ gal}}{46.0\text{ km}}$

d. $\dfrac{50\text{ mg Atenolol}}{1\text{ tablet}}$ and $\dfrac{1\text{ tablet}}{50\text{ mg Atenolol}}$

e. $\dfrac{29\ \mu\text{g}}{1\text{ kg}}$ and $\dfrac{1\text{ kg}}{29\ \mu\text{g}}$

f. $\dfrac{81\text{ mg aspirin}}{1\text{ tablet}}$ and $\dfrac{1\text{ tablet}}{81\text{ mg aspirin}}$

1.55 The unit in the denominator must cancel with the preceding unit in the numerator.

1.57 **a.** 1.75 m **b.** 5.5 L **c.** 1.8 g

1.59 **a.** 473 mL **b.** 65.9 kg **c.** 24 lb **d.** 43 g **e.** 50. lb

1.61 **a.** 66 gal **b.** 3 tablets **c.** 1800 mg

1.63 lead

1.65 **a.** 1.20 g/mL **b.** 4.4 g/mL **c.** 3.10 g/mL **d.** 4.51 g/mL

1.67 **a.** 210 g **b.** 575 g **c.** 62 oz

1.69 **b** and **c**

1.71 **c** and **d**

1.73 **a.** exact **b.** measured **c.** exact **d.** measured

1.75 **a.** length = 6.96 cm, width = 4.75 cm
b. length = 69.6 mm, width = 47.5 mm
c. 3 significant figures
d. 3 significant figures
e. 33.1 cm^2
f. 3 significant figures

1.77 1.8 g/mL

1.79 **a.** 0.000 012 6 L **b.** 3.53×10^2 kg **c.** 125 000 m **d.** 58.7 g **e.** 3.00×10^{-3} s **f.** 0.0108 g

1.81 42 min

1.83 $1.08 per lb

1.85 16 onions

1.87 **a.** 96 crackers **b.** 0.2 oz of fat **c.** 110 g of sodium

1.89 0.35 oz

1.91 185 mg/dL

1.93 790 g

1.95 1200 cm^3

1.97 You would record the mass as 31.075 g. Since the balance will weigh to the nearest 0.001 g, the mass value would be reported to 0.001 g.

1.99 6.4 gal

1.101 0.141 mm

1.103 **a.** 79 cups **b.** 310 cans **c.** 160 tablets

1.105 **a.** 4.3 lb of body fat **b.** 6.2 lb

1.107 A is vegetable oil, B is water, and C is mercury.

Matter and Energy

2

LOOKING AHEAD

"If you've had first aid for a sports injury," says Cort Kim, physical therapist at the Sunrise Sports Medicine Clinic, "you've likely been treated with a cold pack or hot pack. We use them for several kinds of injury. Here, I'm showing how I can use a cold pack to reduce swelling in my patient's shoulder."

A hot or cold pack is just a packaged chemical reaction. When you hit or open the pack to activate it, your action mixes chemicals together and thus initiates the reaction. In a cold pack, the reaction is one that absorbs heat energy, chills the pack, and draws heat from the injury. Hot packs use reactions that release energy, thus warming the pack. In both cases, the reaction proceeds at a moderate pace, so that the pack stays active for a long time and doesn't get too cold or hot.

Every day, we see a variety of materials with many different shapes and forms. To a scientist, all of this material is *matter*. Matter is everywhere around us: the orange juice we had for breakfast, the water we put in the coffee maker, the plastic bag we put our sandwich in, our toothbrush and toothpaste, the oxygen we inhale, and the carbon dioxide we exhale are all forms of matter.

Almost everything we do involves energy. We use energy when we walk, play tennis, study, and breathe. We use energy when we warm water, cook food, turn on lights, use computers, use a washing machine, or drive our cars. Of course, that energy has to come from somewhere. In our bodies, the food we eat provides us with energy. Energy from fossil fuels or the Sun is used to heat a home or water for a pool.

When we look around us, we see that matter takes the physical form of a solid, a liquid, or a gas. Water is a familiar example that we observe in all three states. In the solid state, water can be an ice cube or a snowflake. It is a liquid when it comes out of a faucet or fills a pool. Water forms a gas, or vapor, when it evaporates from wet clothes or boils in a pan. Energy is required to melt ice cubes and to boil water, whereas energy is removed to freeze water in an ice cube tray or to condense water vapor to liquid.

LEARNING GOAL

Classify examples of matter as pure substances or mixtures.

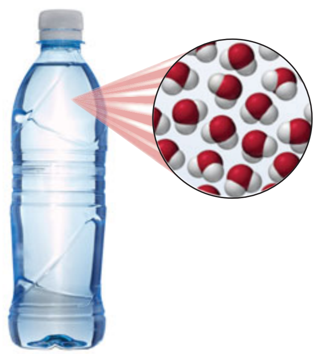
MC TUTORIAL
Classification of Matter

2.1 Classification of Matter

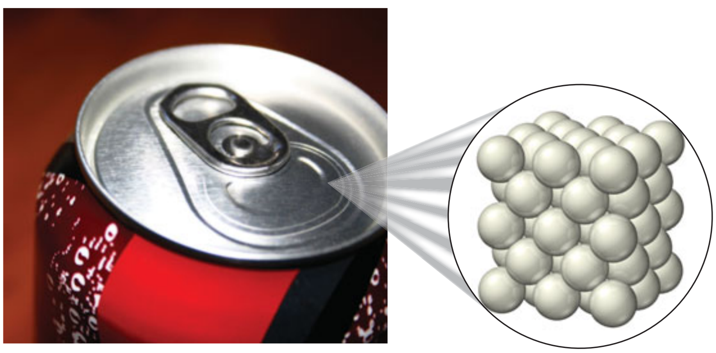
Matter is anything that has mass and occupies space. Matter makes up all things we use such as water, wood, plates, plastic bags, clothes, and shoes. The different types of matter are classified by their composition.

Pure Substances

A **pure substance** is matter that has a fixed or definite composition. There are two kinds of pure substances: elements and compounds. An **element**, the simplest type of a pure substance, is composed of only one type of material such as silver, iron, or aluminum. Every element is composed of *atoms*, which are extremely tiny particles that make up each type of matter. Silver is composed of silver atoms, iron of iron atoms, and aluminum of aluminum atoms. A complete list of the elements is found on the inside front cover of this text.



A molecule of water consists of two atoms of hydrogen (white) for one atom of oxygen (red) and has a formula of H_2O.



An aluminum can consists of many atoms of aluminum.

A **compound** is also a pure substance, but it consists of atoms of two or more elements always chemically combined in the same proportion. In compounds, the atoms are held together by attractions called *bonds*, which form small groups of atoms called molecules. For example, a molecule of the compound water has two hydrogen atoms for every one oxygen atom and is represented by the formula H_2O. The compound hydrogen peroxide is also a combination of hydrogen and oxygen, but it has two hydrogen atoms for every two oxygen atoms and is represented by the formula H_2O_2. Water (H_2O) and hydrogen peroxide (H_2O_2) are different compounds, which means they have different properties.

An important difference between compounds and elements is that compounds can be broken down by chemical processes into simpler substances, whereas elements cannot be broken down further. For example, ordinary table salt consists of the compound NaCl, which can be separated by chemical processes into sodium metal and chlorine gas, as seen in Figure 2.1. However, compounds such as NaCl cannot be separated into simpler substances by using physical methods such as boiling or sifting.

A molecule of hydrogen peroxide consists of two atoms of hydrogen (white) for two atoms of oxygen (red) and has a formula of H_2O_2.

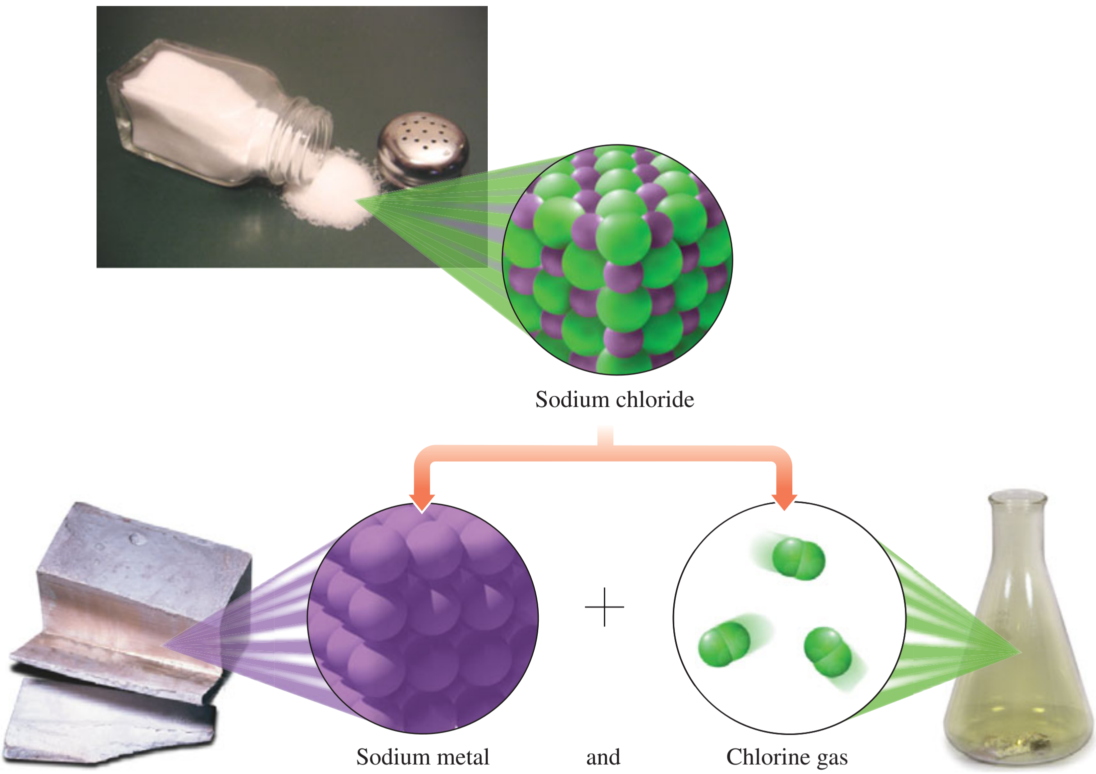


FIGURE 2.1 The decomposition of salt, NaCl, produces the elements sodium and chlorine.
Q How do elements and compounds differ?

Mixtures

Much of the matter in our everyday lives consists of mixtures. In a **mixture**, two or more substances are physically mixed, but not chemically combined. The air we breathe is a mixture of mostly oxygen and nitrogen gases. The steel in buildings and railroad tracks is a mixture of iron, nickel, carbon, and chromium. The brass in doorknobs and fixtures is a mixture of zinc and copper. Tea, coffee, and ocean water are mixtures too. In any mixture, the proportions of the components can vary. For example, two sugar–water mixtures may look the same, but the one with the higher ratio of sugar to water would taste sweeter. Different types of brass have different properties such as color or strength depending on the ratio of copper to zinc (see Figure 2.2).

Physical processes can be used to separate mixtures because there are no chemical interactions between the components. For example, different coins such as nickels,

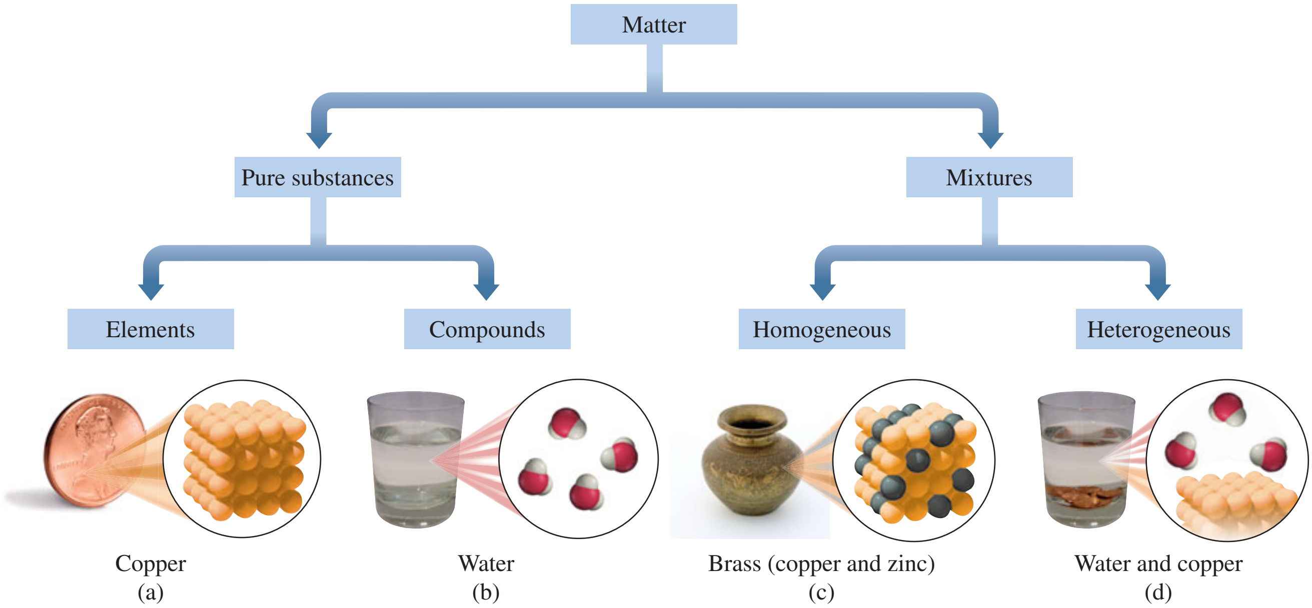


FIGURE 2.2 Matter is organized by its components: elements, compounds, and mixtures. (a) The element copper consists of copper atoms. (b) The compound water consists of H_2O molecules. (c) Brass is a homogeneous mixture of copper and zinc atoms. (d) Copper metal in water is a heterogeneous mixture of Cu atoms and H_2O molecules.
Q Why are copper and water pure substances, but brass is a mixture?

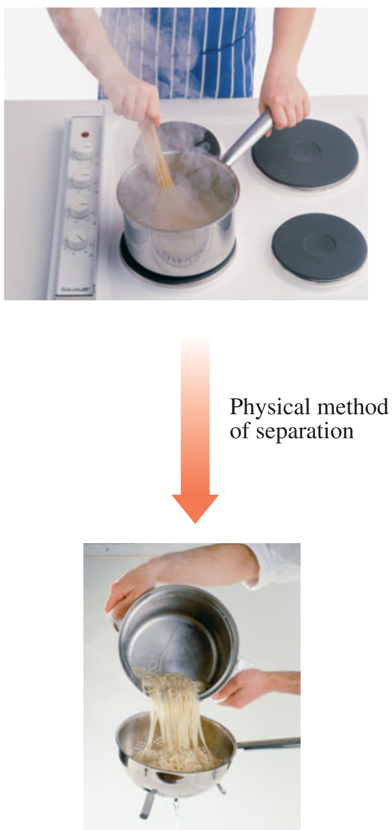


FIGURE 2.3 A mixture of spaghetti and water is separated using a strainer, a physical method of separation.
Q Why can physical methods be used to separate mixtures but not compounds?

dimes, and quarters can be separated by size; iron particles mixed with sand can be picked up with a magnet; and water is separated from cooked spaghetti by using a strainer (see Figure 2.3).

CONCEPT CHECK 2.1

Pure Substances and Mixtures

Classify each of the following as a pure substance or a mixture:

a. sugar in a sugar bowl
b. a collection of nickels and dimes
c. coffee with milk and sugar

ANSWER

a. Sugar is a compound, which is a pure substance.
b. The nickels and dimes are physically mixed, but not chemically combined, which makes the collection a mixture.
c. The coffee, milk, and sugar are physically mixed, but not chemically combined, which makes it a mixture.

Types of Mixtures

Mixtures are classified further as homogeneous or heterogeneous. In a *homogeneous mixture*, also called a *solution*, the composition is uniform throughout the sample. Familiar examples of homogeneous mixtures are air, which contains oxygen and nitrogen gases, and sea water, a solution of salt and water.

In a *heterogeneous mixture*, the components do not have a uniform composition throughout the sample. For example, a mixture of oil and water is heterogeneous because the oil floats on the surface of the water. Other examples of heterogeneous mixtures include the raisins in a cookie and the bubbles in a soda.

In the chemistry laboratory, mixtures are separated by various methods. Solids are separated from liquids by *filtration*, which involves pouring a mixture through a filter paper set in a funnel. In *chromatography*, different components of a liquid mixture separate as they move at different rates up the surface of a piece of chromatography paper.

A mixture of a liquid and a solid is separated by filtration.

SAMPLE PROBLEM 2.1

Classifying Mixtures

Classify each of the following as a pure substance (element or compound) or a mixture (homogeneous or heterogeneous):

a. copper in copper wire
b. a chocolate chip cookie
c. Nitrox, a breathing mixture of oxygen and nitrogen for scuba diving

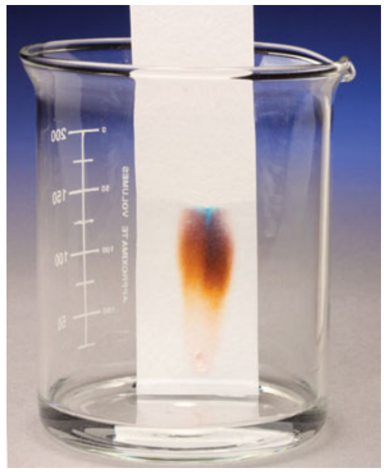
SOLUTION

a. Copper is an element, which is a pure substance.
b. A chocolate chip cookie does not have a uniform composition, which makes it a heterogeneous mixture.
c. The gases oxygen and nitrogen have a uniform composition in Nitrox, which makes it a homogeneous mixture.

STUDY CHECK 2.1

A salad dressing is prepared with oil, vinegar, and chunks of blue cheese. Is this a homogeneous or heterogeneous mixture?

Oil and water form a heterogeneous mixture.



Different substances are separated as they travel at different rates up the surface of chromatography paper.

Chemistry Link to Health

BREATHING MIXTURES FOR SCUBA

The air we breathe is composed mostly of the gases oxygen (21%) and nitrogen (79%). The homogeneous breathing mixtures used by scuba divers differ from the air we breathe depending on the depth of the dive. Nitrox is a mixture of oxygen and nitrogen, but with more oxygen gas (up to 32%) and less nitrogen gas (68%) than air. A breathing mixture with less nitrogen gas decreases the risk of nitrogen narcosis associated with breathing regular air while diving. Heliox is a breathing mixture of oxygen and helium gases typically used for diving to more than 200 feet. With deep dives, there is more chance of nitrogen narcosis, but by replacing nitrogen with helium, it does not occur. However, at dive depths over 300 ft, helium is associated with severe shaking and body temperature drop.

A breathing mixture used for dives over 400 ft is Trimix, which contains oxygen, helium, and some nitrogen. The addition of nitrogen lessens the problem of shaking that comes with breathing high levels of helium. Both Heliox and Trimix are used by only professional, military, or other highly trained divers.

A Nitrox mixture is used to fill scuba tanks.

QUESTIONS AND PROBLEMS

Classification of Matter

2.1 Classify each of the following as a pure substance or a mixture:
a. baking soda ($NaHCO_3$) **b.** a blueberry muffin
c. ice (H_2O) **d.** zinc (Zn)
e. Trimix (oxygen, nitrogen, and helium) in a scuba tank

2.2 Classify each of the following as a pure substance or a mixture:
a. a soft drink **b.** propane (C_3H_8)
c. a cheese sandwich **d.** an iron (Fe) nail
e. salt substitute (KCl)

2.3 Classify each of the following pure substances as an element or a compound:
a. a silicon (Si) chip **b.** hydrogen peroxide (H_2O_2)
c. oxygen (O_2) **d.** rust (Fe_2O_3)
e. methane (CH_4) in natural gas

2.4 Classify each of the following pure substances as an element or a compound:
a. helium gas (He) **b.** mercury (Hg) in a thermometer
c. sugar ($C_{12}H_{22}O_{11}$) **d.** sulfur (S)
e. lye (NaOH)

2.5 Classify each of the following mixtures as homogeneous or heterogeneous:
a. vegetable soup **b.** sea water
c. tea **d.** tea with ice and lemon slices
e. fruit salad

2.6 Classify each of the following mixtures as homogeneous or heterogeneous:
a. nonfat milk **b.** chocolate-chip ice cream
c. gasoline **d.** peanut-butter-and-jelly sandwich
e. cranberry juice

LEARNING GOAL

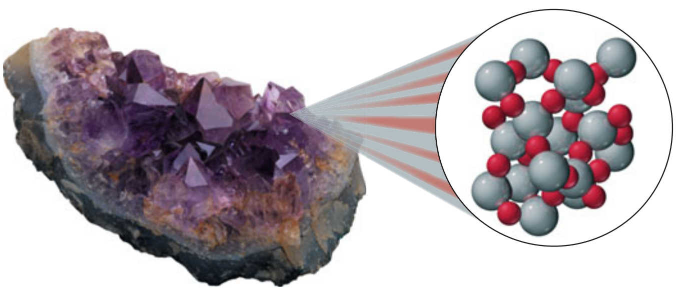
Identify the states and physical and chemical properties of matter.

2.2 States and Properties of Matter

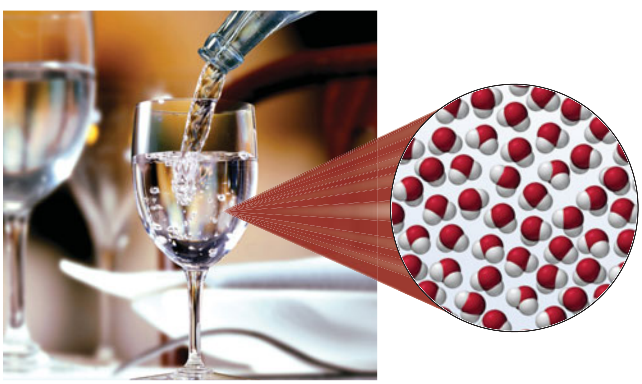
On Earth, matter exists in one of three *physical forms* called the **states of matter**: *solids*, *liquids*, and *gases*. A **solid**, such as a pebble or a baseball, has a definite shape and volume. You can probably recognize several solids within your reach right now such as books, pencils, or a computer mouse. In a *solid*, strong attractive forces hold the particles such as atoms or molecules close together. The particles are arranged in such a rigid pattern they can only vibrate slowly in fixed positions. For many solids, this rigid structure produces a crystal such as that seen in amethyst.

A **liquid** has a definite volume, but not a definite shape. In a *liquid*, the particles move in random directions but are sufficiently attracted to each other to maintain a definite volume, although not a rigid structure. Thus, when water, oil, or vinegar is poured from one container to another, the liquid maintains its own volume but takes the shape of the new container.

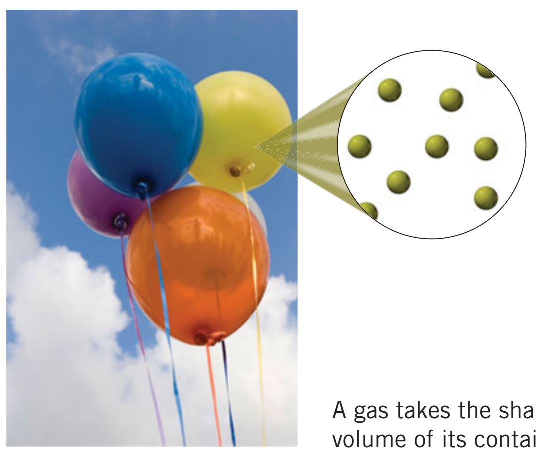
A **gas** does not have a definite shape or volume. In a *gas*, the particles are far apart, have little attraction to each other, and move at high speeds, taking the shape and volume of their container. When you inflate a bicycle tire, the air, which is a gas, fills the entire volume of the tire. The propane gas in a tank fills the entire volume of the tank. Table 2.1 compares the three states of matter.



Amethyst, a solid, is a purple form of quartz (SiO_2).



Water as a liquid takes the shape of its container.



A gas takes the shape and volume of its container.

TABLE 2.1 A Comparison of Solids, Liquids, and Gases

Characteristic	Solid	Liquid	Gas
Shape	Has a definite shape	Takes the shape of the container	Takes the shape of the container
Volume	Has a definite volume	Has a definite volume	Fills the volume of the container
Arrangement of particles	Fixed, very close	Random, close	Random, far apart
Interaction between particles	Very strong	Strong	Essentially none
Movement of particles	Very slow	Moderate	Very fast
Examples	Ice, salt, iron	Water, oil, vinegar	Water vapor, helium, air

Physical Properties and Physical Changes

One way to describe matter is to observe its properties. For example, if you were asked to describe yourself, you might list your characteristics such as the color of your eyes and skin or the length, color, and texture of your hair.

Physical properties are those characteristics that can be observed or measured without affecting the identity of a substance. In chemistry, typical physical properties include the shape, color, melting point, boiling point, and physical state of a substance. For example, a penny has the physical properties of a round shape, an orange-red color, solid state, and a shiny luster. Table 2.2 gives more examples of physical properties of copper found in pennies, electrical wiring, and copper pans.

Water is a substance that is commonly found in all three states: solid, liquid, and gas. When matter undergoes a **physical change**, its state or its appearance will change, but its composition remains the same. The solid state of water, snow or ice, has a different appearance than its liquid or gaseous state, but all three states are water (see Figure 2.4).

The physical appearance of a substance can change in other ways too. Suppose that you dissolve some salt in water. The appearance of the salt changes, but you could re-form the salt crystals by heating the mixture and evaporating the water. Thus in a physical change of state, no new substances are produced. Table 2.3 gives more examples of physical changes.

TUTORIAL
Properties and Changes of Matter

Copper, used in cookware, is a good conductor of heat.

TABLE 2.2 Some Physical Properties of Copper

State at 25 °C	Solid
Color	Orange-red
Odor	Odorless
Melting point	1083 °C
Boiling point	2567 °C
Luster	Shiny
Conduction of electricity	Excellent
Conduction of heat	Excellent

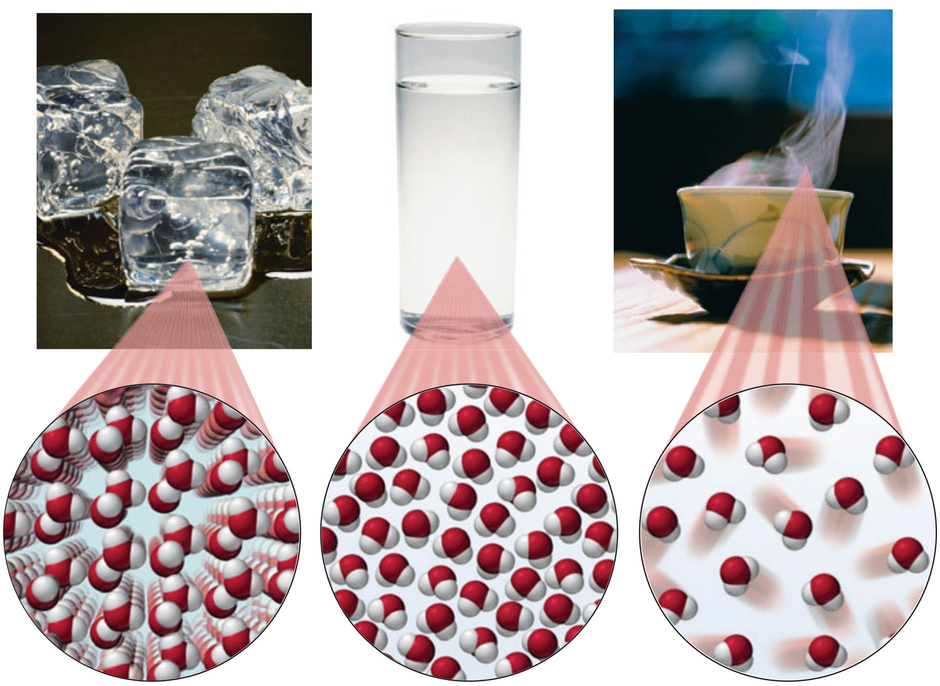


FIGURE 2.4 Water exists as a solid, as a liquid, or as a gas.

Q In what state of matter does water have a definite volume, but not a definite shape?

In a physical change, a gold ingot is hammered to form gold leaf.

TABLE 2.3 Examples of Some Physical Changes

Type of Physical Change	Example
Change of state	Water boiling Freezing of liquid water to solid water (ice)
Change of appearance	Dissolving sugar in water
Change of shape	Hammering a gold ingot into shiny gold leaf Drawing copper into thin copper wire
Change of size	Cutting paper into tiny pieces for confetti Grinding pepper into smaller particles

CONCEPT CHECK 2.2

States of Matter

Identify the state(s) of matter described by the substance in each of the following:

a. Its volume does not change in a different container.
b. Its shape depends on the container.
c. It has a definite shape and volume.

ANSWER

a. Both a solid and a liquid have their own volume that does not depend on the volume of their container.
b. Both a liquid and a gas take the shape of their containers.
c. A solid has a rigid arrangement of particles that gives it a definite shape and volume.

Chemical Properties and Chemical Changes

Chemical properties are those that describe the ability of a substance to change into a new substance. When a **chemical change** takes place, the original substance is converted into one or more new substances, which have different physical and chemical properties. For example, methane (CH_4) in natural gas can burn because it has the chemical property of being flammable. When methane burns in oxygen (O_2), it is converted to water (H_2O) and carbon dioxide (CO_2), which have different physical and chemical properties than the methane and oxygen. Rusting or corrosion is a chemical property of iron. In the rain, an iron nail undergoes a chemical change when it reacts with oxygen in the air to form rust (Fe_2O_3), a new substance. Table 2.4 gives some examples of chemical changes.


Flan has a topping of caramelized sugar.

TABLE 2.4 Examples of Some Chemical Changes

Type of Chemical Change	Change in Chemical Properties
Silver tarnishes	Shiny, silver metal reacts in air to give a black, grainy coating.
Methane burns	Methane burns with a bright flame, producing water vapor and carbon dioxide.
Sugar caramelizes	At high temperatures, white, granular sugar changes to a smooth, caramel-colored substance.
Iron rusts	Iron, which is gray and shiny, combines with oxygen to form orange-red rust.

CONCEPT CHECK 2.3

Physical and Chemical Properties

Classify each of the following as a physical or chemical property:

a. Gasoline is a liquid at room temperature.
b. Gasoline burns in air.
c. Gasoline has a pungent odor.

ANSWER

a. A liquid is a state of matter, which makes it a physical property.
b. When gasoline burns, it changes to different substances with new properties, which is a chemical property.
c. The odor of gasoline is a physical property.

SAMPLE PROBLEM 2.2

Physical and Chemical Changes

Classify each of the following as a physical or chemical change:

a. An ice cube melts to form liquid water.
b. An enzyme breaks down the lactose in milk.
c. Garlic is chopped into small pieces.

SOLUTION

a. A physical change occurs when the ice cube changes state from solid to liquid.
b. A chemical change occurs when an enzyme breaks down lactose into simpler substances.
c. A physical change occurs when the size of an object changes.

STUDY CHECK 2.2

Which of the following are chemical changes?

a. Water freezes on a pond.
b. Gas bubbles form when baking powder is placed in vinegar.
c. A log is chopped for firewood.
d. A log burns in a fireplace.

QUESTIONS AND PROBLEMS

States and Properties of Matter

2.7 Indicate whether each of the following describes a gas, a liquid, or a solid:
a. This substance has no definite volume or shape.
b. The particles in a substance do not interact with each other.
c. The particles in a substance are held in a rigid structure.

2.8 Indicate whether each of the following describes a gas, a liquid, or a solid:
a. This substance has a definite volume but takes the shape of its container.
b. The particles in a substance are very far apart.
c. This substance occupies the entire volume of the container.

2.9 Describe each of the following as a physical or chemical property:
a. Chromium is a steel-gray solid.
b. Hydrogen reacts readily with oxygen.
c. Nitrogen freezes at -210 °C.
d. Milk will sour when left in a warm room.
e. Butane gas in an igniter burns in oxygen.

2.10 Describe each of the following as a physical or chemical property:
a. Neon is a colorless gas at room temperature.
b. Apple slices turn brown when they are exposed to air.
c. Phosphorus will ignite when exposed to air.
d. At room temperature, mercury is a liquid.
e. Propane gas is compressed to a liquid for placement in a small cylinder.

2.11 What type of change, physical or chemical, takes place in each of the following?
a. Water vapor condenses to form rain.
b. Cesium metal reacts explosively with water.
c. Gold melts at 1064 °C.
d. A puzzle is cut into 1000 pieces.
e. Sugar dissolves in water.

2.12 What type of change, physical or chemical, takes place in each of the following?
a. Gold is hammered into thin sheets.
b. A silver pin tarnishes in the air.
c. A tree is cut into boards at a saw mill.
d. Food is digested.
e. A chocolate bar melts.

2.13 Describe each property of the element fluorine as physical or chemical.
a. is highly reactive
b. is a gas at room temperature
c. has a pale, yellow color
d. will explode in the presence of hydrogen
e. has a melting point of −220 °C

2.14 Describe each property of the element zirconium as physical or chemical.
a. melts at 1852 °C
b. is resistant to corrosion
c. has a grayish-white color
d. ignites spontaneously in air when finely divided
e. is a shiny metal

LEARNING GOAL

Identify energy as potential or kinetic; convert between units of energy.

2.3 Energy

When you are running, walking, dancing, or thinking, you are using energy to do *work*, which is any activity that requires energy. In fact, **energy** is defined as the ability to do work. Suppose you are climbing a steep hill and you become too tired to go on. At that moment, you do not have the energy to do any more work. Now suppose you sit down and have lunch. In a while, you will have obtained some energy from the food, and you will be able to do more work and complete the climb (see Figure 2.5).

FIGURE 2.5 Work is done as the rock climber moves up the cliff. At the top, the climber has more potential energy than when she started the climb.
Q What happens to the potential energy of the climber as she descends?

Potential and Kinetic Energy

Energy can be classified as potential energy or kinetic energy. **Kinetic energy** is the energy of motion. Any object that is moving has kinetic energy. **Potential energy** is determined by the position of an object or by the chemical composition of a substance. A boulder resting on top of a mountain has potential energy because of its location. If the boulder rolls down the mountain, the potential energy becomes kinetic energy. Water stored in a reservoir has potential energy. When the water goes over the dam and falls to the stream below, its potential energy becomes kinetic energy. Foods and fossil fuels have potential energy in their molecules. When you digest food or burn gasoline in your car, potential energy is converted to kinetic energy to do work.

When water flows from the top of a dam, potential energy is converted to kinetic energy.

CONCEPT CHECK 2.4

Potential and Kinetic Energy

Identify each of the following as an example of potential or kinetic energy:

a. gasoline
b. skating
c. a candy bar

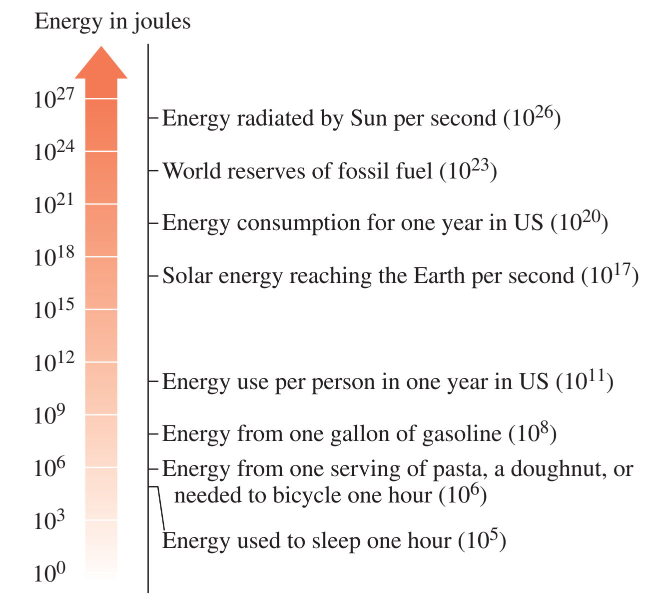
ANSWER

a. Gasoline is burned to provide energy and heat; it contains potential energy (stored) in its molecules.
b. A skater uses energy to move; skating is kinetic energy (energy of motion).
c. A candy bar has potential energy. When digested, it provides energy for the body to do work.

Heat and Units of Energy

Heat, also known as *thermal energy*, is associated with the motion of particles. A frozen pizza feels cold because heat flows from your hand into the pizza. The faster the particles move, the greater the heat or thermal energy of the substance. In the frozen pizza, the particles are moving very slowly. As heat is added and the pizza becomes warmer, the motions of the particles in the pizza increase. Eventually, the particles have enough energy to make the pizza hot and ready to eat.

TABLE 2.5 A Comparison of Energy for Various Resources



Units of Energy

The SI unit of energy and work is the **joule (J)** (pronounced "jewel"). The joule is a small amount of energy, so scientists often use the kilojoule (kJ), 1000 joules. To heat water for one cup of tea, you need about 75 000 J or 75 kJ of heat. Table 2.5 shows a comparison of energy in joules for several energy sources.

You may be more familiar with the unit **calorie (cal)**, from the Latin *calor*, meaning "heat." The calorie was originally defined as the amount of energy (heat) needed to raise the temperature of 1 g of water by 1 °C. Now one calorie is defined as *exactly* 4.184 J. This equality can be written as two conversion factors:

$$1 \text{ cal} = 4.184 \text{ J (exact)}$$

$$\frac{4.184 \text{ J}}{1 \text{ cal}} \quad \text{and} \quad \frac{1 \text{ cal}}{4.184 \text{ J}}$$

MC TUTORIAL Heat

TUTORIAL Energy Conversions

One *kilocalorie* (kcal) is equal to 1000 calories, and one *kilojoule* (kJ) is equal to 1000 joules.

$$1 \text{ kcal} = 1000 \text{ cal}$$

$$1 \text{ kJ} = 1000 \text{ J}$$

SAMPLE PROBLEM 2.3

Energy Units

When 1.0 g of diesel burns in an diesel car engine, 48 000 J are released. What is this quantity of energy in calories?

SOLUTION

Step 1 **Given** 48 000 J **Need** calories (cal)

Step 2 **Plan**

joules → Energy factor → calories

Step 3 **Equalities/Conversion Factors**

$$1 \text{ cal} = 4.184 \text{ J}$$

$$\frac{1 \text{ cal}}{4.184 \text{ J}} \quad \text{and} \quad \frac{4.184 \text{ J}}{1 \text{ cal}}$$

Diesel fuel reacts in a car engine to produce energy.

Step 4 Set Up Problem

$$\underset{\text{Two SFs}}{48\,000\,\cancel{J}} \times \underset{\text{Exact}}{\boxed{\frac{1\text{ cal}}{4.184\,\cancel{J}}}} = \underset{\text{Two SFs}}{11\,000\text{ cal }(1.1 \times 10^4\text{ cal})}$$

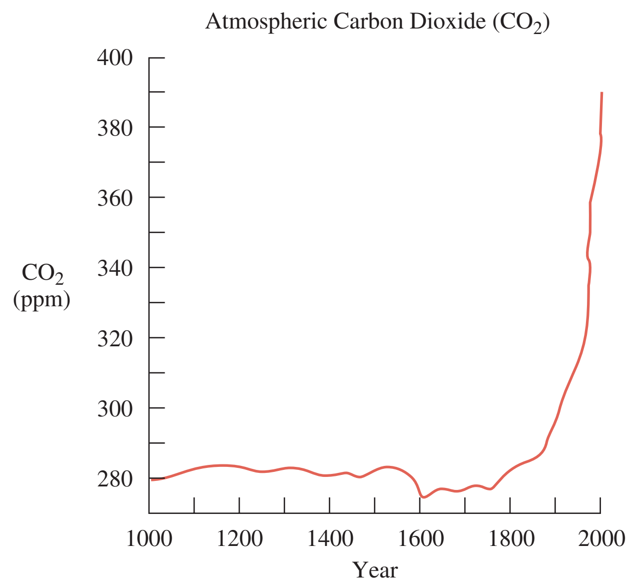
STUDY CHECK 2.3

The burning of 1.0 g of coal produces 8.4 kcal. How many joules are produced?

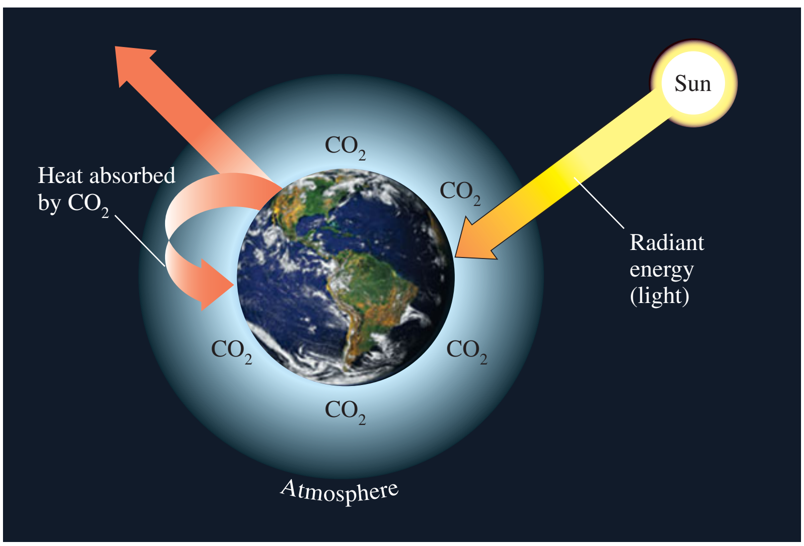
Chemistry Link to the Environment

CARBON DIOXIDE AND GLOBAL WARMING

Earth's climate is a product of interactions between sunlight, the atmosphere, and the oceans. The Sun provides us with energy in the form of solar radiation. Some of this radiation is reflected back into space. The rest is absorbed by the clouds, atmospheric gases including carbon dioxide, and Earth's surface. For millions of years, concentrations of carbon dioxide (CO_2) have fluctuated. However in the last 100 years, the amount of carbon dioxide (CO_2) gas in our atmosphere has increased significantly. From the years 1000 to 1800, the atmospheric carbon dioxide averaged 280 ppm. But since the beginning of the Industrial Revolution in 1800 up until 2005, the level of atmospheric carbon dioxide has risen from about 280 ppm to about 390 ppm, a 40% increase.



As the atmospheric CO_2 level increases, more solar radiation is trapped by the atmospheric gases, which raises the temperature at the surface of Earth. Some scientists have estimated that if the carbon dioxide level doubles from its level before the Industrial Revolution, the average global temperature could increase by 2 to 4.4 °C. Although this seems to be a small temperature change, it could have dramatic impact worldwide. Even now, glaciers and snow cover in much of the world have diminished. Ice sheets in Antarctica and Greenland are melting faster and breaking apart. Although no one knows for sure how rapidly the ice in the polar regions is melting, this accelerating change will contribute to a rise in sea level. In the twentieth century, the sea level increased by about 20 cm. Some scientists predict the sea level will rise 1 m in this century. Such an increase will have a major impact on coastal areas.



Heat from the Sun is trapped by the CO_2 layer in the atmosphere.

Until recently, the carbon dioxide level was maintained as algae in the oceans and the trees in the forests utilized the carbon dioxide. However, the ability of these and other forms of plant life to absorb carbon dioxide is not keeping up with the increase in carbon dioxide. Most scientists agree that the primary source of the increase of carbon dioxide is the burning of fossil fuels such as gasoline, coal, and natural gas. The cutting and burning of trees in the rain forests (deforestation) also reduces the amount of carbon dioxide removed from the atmosphere.

Worldwide efforts are being made to reduce the carbon dioxide produced by burning fossil fuels that heat our homes, run our cars, and provide energy for industries. Scientists are exploring ways to provide alternative energy sources and to reduce the effects of deforestation. Meanwhile, we can reduce energy use in our homes by using appliances that are more energy efficient such as replacing incandescent light bulbs with fluorescent lights. Such an effort worldwide will reduce the possible impact of global warming and at the same time save our fuel resources.

QUESTIONS AND PROBLEMS

Energy

2.15 Discuss the changes in the potential and kinetic energy of a roller-coaster ride as the roller-coaster car climbs to the top and goes down the other side.

2.16 Discuss the changes in the potential and kinetic energy of a ski jumper taking the elevator to the top of the jump and going down the ramp.

2.17 Indicate whether each of the following statements describes potential or kinetic energy:
a. water at the top of a waterfall
b. kicking a ball
c. the energy in a lump of coal
d. a skier at the top of a hill

2.18 Indicate whether each of the following statements describes potential or kinetic energy:
a. the energy in your food
b. a tightly wound spring
c. an earthquake
d. a car speeding down the freeway

2.19 The energy needed to keep a 75-watt lightbulb burning for 1.0 h is 270 kJ. Calculate the energy required to keep the lightbulb burning for 3.0 h in each of the following energy units:
a. joules **b.** kilocalories

2.20 A person uses 750 kcal on a long walk. Calculate the energy used for the walk in each of the following energy units:
a. joules **b.** kilojoules

2.4 Temperature

LEARNING GOAL

Given a temperature, calculate a corresponding temperature on another scale.

Temperatures in science, and in most of the world, are measured and reported in *Celsius* (°C) units. On the Celsius scale, the reference points are the freezing point of water, 0 °C, and the boiling point, 100 °C. In the United States, everyday temperatures are commonly reported in *Fahrenheit* (°F) units. On the Fahrenheit scale, water freezes at 32 °F and boils at 212 °F. A typical room temperature of 22 °C would be the same as 72 °F. A normal body temperature of 37.0 °C is 98.6 °F.

On the Celsius and Fahrenheit temperature scales, the temperature difference between freezing and boiling is divided into smaller units called *degrees*. The Celsius scale has 100 degrees between the freezing and boiling points of water whereas the Fahrenheit scale has 180 degrees between the freezing and boiling points. That makes a degree Celsius almost twice the size of a Fahrenheit degree: 1 °C = 1.8 °F (see Figure 2.6).

TUTORIAL
Temperature Conversions

$$180\ \text{Fahrenheit degrees} = 100\ \text{degrees Celsius}$$

$$\frac{180\ \text{Fahrenheit degrees}}{100\ \text{degrees Celsius}} = \frac{1.8\ °\text{F}}{1\ °\text{C}}$$

A digital ear thermometer is used to measure body temperature.

We can write a temperature equation that relates a Fahrenheit temperature and its corresponding Celsius temperature.

$$T_F = \underbrace{1.8(T_C)}_{\text{Changes °C to °F}} + \underbrace{32}_{\text{Adjusts freezing point}}$$

In this equation, the Celsius temperature is multiplied by 1.8 to change °C to °F; then 32 is added to adjust the freezing point from 0 °C to 32 °F. Both values, 1.8 and 32, are exact numbers.

To convert from Fahrenheit to Celsius, the temperature equation is rearranged for T_C.

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

Scientists have learned that the coldest temperature possible is −273 °C (more precisely, −273.15 °C). On the *Kelvin* scale, this temperature, called *absolute zero*, has

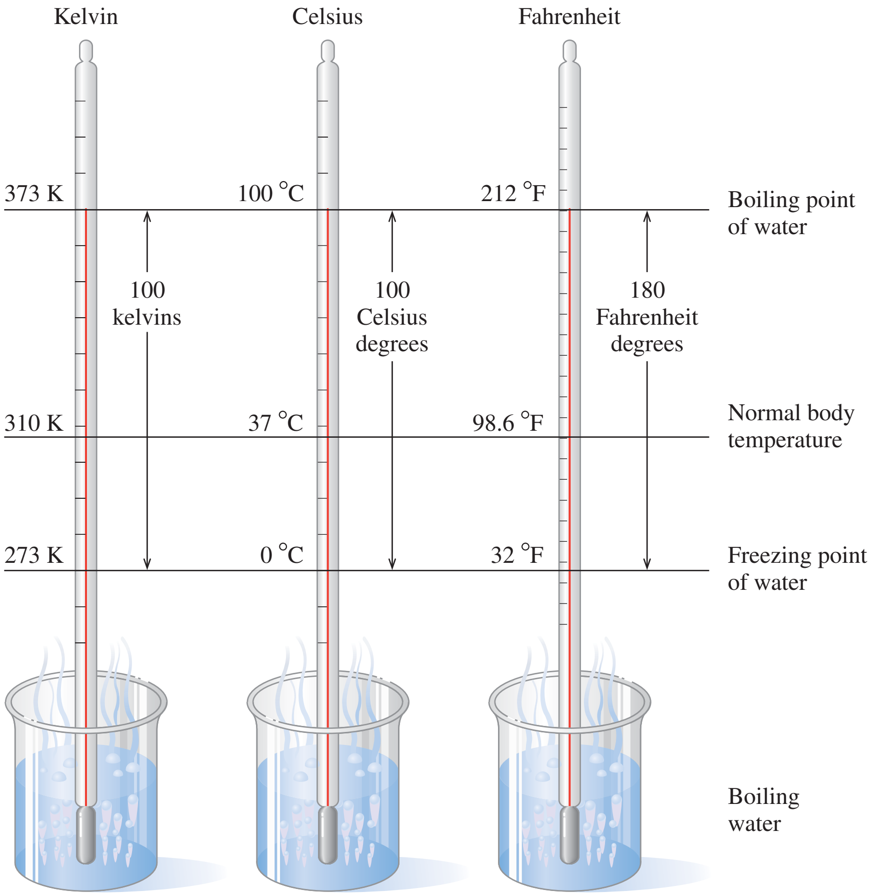


FIGURE 2.6 A comparison of the Fahrenheit, Celsius, and Kelvin temperature scales between the freezing and boiling points of water.

Q What is the difference in the values for freezing on the Celsius and Fahrenheit temperature scales?

the value of 0 K. Units on the Kelvin scale are called kelvins (K); no degree symbol is used. Because there are no lower temperatures, the Kelvin scale has no negative temperature values. Between the freezing point of water, 273 K, and the boiling point, 373 K, there are 100 kelvins, which makes a kelvin equal in size to a Celsius degree.

$$1\ \text{K} = 1\ ^\circ\text{C}$$

We can write an equation that relates a Celsius temperature to its corresponding Kelvin temperature by adding 273. Table 2.6 gives a comparison of some temperatures on the three scales.

$$T_K = T_C + 273$$

TABLE 2.6 A Comparison of Temperatures

Example	Fahrenheit (°F)	Celsius (°C)	Kelvin (K)
Sun	9937	5503	5776
A hot oven	450	232	505
A desert	120	49	322
A high fever	104	40	313
Room temperature	70	21	294
Water freezes	32	0	273
A northern winter	−66	−54	219
Helium boils	−452	−269	4
Absolute zero	−459	−273	0

SAMPLE PROBLEM 2.4

Converting Celsius to Fahrenheit

A room is heated to 22 °C. If that temperature is lowered by 1 °C, it can save as much as 5% in energy costs. What temperature, in Fahrenheit degrees, should be set to lower the temperature by 1 °C?

SOLUTION

Step 1 **Given** 22 °C − 1 °C = 21 °C **Need** T_F

Step 2 **Plan**

T_C → Temperature equation → T_F

Step 3 **Equality/Conversion Factor**

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

Step 4 **Set Up Problem** Substitute the Celsius temperature into the equation and solve.

$$T_F = \underbrace{1.8}_{\text{Exact}}(\underbrace{21}_{\text{Two SFs}}) + 32$$

$$T_F = 38 + 32 \quad \text{1.8 is exact; 32 is exact}$$

$$= 70.\ °\text{F} \quad \text{Answer to the ones place}$$

In the equation, *the values of 1.8 and 32 are exact numbers*, which do not affect the number of SFs.

STUDY CHECK 2.4

In the process of making ice cream, rock salt is added to crushed ice to chill the ice cream mixture. If the temperature drops to −11 °C, what is it in Fahrenheit degrees?

SAMPLE PROBLEM 2.5

Converting Fahrenheit to Celsius

In a type of cancer treatment called *thermotherapy*, temperatures as high as 113 °F are used to destroy cancer cells. What is that temperature in degrees Celsius?

SOLUTION

Step 1 **Given** 113 °F **Need** T_C

Step 2 **Plan**

T_F → Temperature equation → T_C

Step 3 **Equality/Conversion Factor**

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

Step 4 **Set Up Problem** Substitute the Fahrenheit temperature into the equation and solve.

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

Career Focus

SURGICAL TECHNOLOGIST

"As a surgical technologist, I assist the doctors during surgeries," says Christopher Ayars, surgical technologist, Kaiser Hospital. "I am there to help during general or orthopedic surgery by passing instruments, holding retractors, and maintaining the sterile field. Our equipment for surgery is sterilized by steam that is heated to 270 °F, which is the same as 130 °C."

Surgical technologists assist with surgical procedures by preparing and maintaining surgical equipment, instruments, and supplies; providing patient care in an operating room setting; preparing and maintaining a sterile field; and ensuring that there are no breaks in aseptic technique. Instruments, which have been sterilized, are wrapped and sent to surgery where they are checked again before they are opened.

$$T_C = \frac{(113 - 32)}{1.8}$$ 32 is exact; 1.8 is exact

$$= \frac{81}{1.8} = 45\ °C$$ Answer to the ones place

STUDY CHECK 2.5

A child has a temperature of 103.6 °F. What is this temperature on a Celsius thermometer?

SAMPLE PROBLEM 2.6

Converting from Celsius to Kelvin Temperature

A dermatologist may use liquid cryogenic nitrogen at −196 °C to remove skin lesions and some skin cancers. What is the temperature of the liquid nitrogen in kelvins?

SOLUTION

Step 1 **Given** −196 °C **Need** T_K

Step 2 **Plan**

T_C → Temperature equation → T_K

Step 3 **Equality/Conversion Factor**

$$T_K = T_C + 273$$

Step 4 **Set Up Problem** Substitute the Celsius temperature into the equation and solve.

$$T_K = T_C + 273$$

$$T_K = -196 + 273$$

$$= 77\ K$$ Answer to the ones place

STUDY CHECK 2.6

On the planet Mercury, the average night temperature is 13 K, and the average day temperature is 683 K. What are these temperatures in degrees Celsius?

QUESTIONS AND PROBLEMS

Temperature

2.21 Your friend who is visiting from Canada just took her temperature. When she reads 99.8 °F, she becomes concerned that she is quite ill. How would you explain this temperature to your friend?

2.22 You have a friend who is using a recipe for flan from a Mexican cookbook. You notice that he set your oven temperature at 175 °F. What would you advise him to do?

2.23 Solve the following temperature conversions:
a. 37.0 °C = ________ °F **b.** 65.3 °F = ________ °C
c. −27 °C = ________ K **d.** 62 °C = ________ K
e. 114 °F = ________ °C

2.24 Solve the following temperature conversions:
a. 25 °C = ________ °F **b.** 155 °C = ________ °F
c. −25 °F = ________ °C **d.** 224 K = ________ °C
e. 145 °C = ________ K

2.25 **a.** A patient with hyperthermia has a temperature of 106 °F. What does this read on a Celsius thermometer?
b. Because high fevers can cause convulsions in children, the doctor wants to be called if the child's temperature goes over 40.0 °C. Should the doctor be called if a child has a temperature of 103 °F?

2.26 **a.** Hot compresses for a patient are prepared with water heated to 145 °F. What is the temperature of the hot water in degrees Celsius?
b. During extreme hypothermia, a boy's temperature dropped to 20.6 °C. What was his temperature on the Fahrenheit scale?

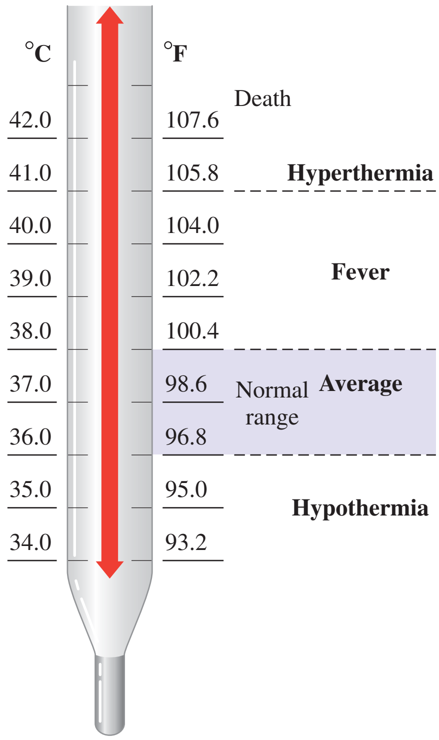
Chemistry Link to Health

VARIATION IN BODY TEMPERATURE

Normal body temperature is considered to be 37.0 °C, although it varies throughout the day and from person to person. Oral temperatures of 36.1 °C are common in the morning and climb to a high of 37.2 °C between 6 P.M. and 10 P.M. Temperatures above 37.2 °C for a person at rest are usually an indication of illness. Individuals who are involved in prolonged exercise may also experience elevated temperatures. Body temperatures of marathon runners can range from 39 °C to 41 °C as heat production during exercise exceeds the body's ability to lose heat.

Changes of more than 3.5 °C from the normal body temperature begin to interfere with bodily functions. At body temperatures above 41 °C, a condition called *hyperthermia* or heat stroke may occur in which sweat production stops, the pulse rate is elevated, and respiration becomes weak and rapid. A person with hyperthermia may become lethargic and lapse into a coma. In children, convulsions can occur, which may lead to permanent brain damage. Damage to internal organs is a major concern, and treatment, which must be immediate, may include immersing the person in an ice-water bath.

At the low temperature extreme of *hypothermia*, body temperature can drop as low as 28.5 °C. The person may appear cold and pale and have an irregular heartbeat. Unconsciousness can occur if the body temperature drops below 26.7 °C. Respiration becomes slow and shallow, and oxygenation of the tissues decreases. Treatment involves providing oxygen and increasing blood volume with glucose and saline fluids. Injecting warm fluids (37.0 °C) into the peritoneal cavity may restore the internal temperature.



2.5 Specific Heat

LEARNING GOAL

Use specific heat to calculate heat loss or gain, temperature change, or mass of a sample.

TUTORIAL
Heat Capacity

Every substance can absorb heat. When you bake a potato, you place it in a hot oven. If you are cooking pasta, you add the pasta to boiling water. You already know that adding heat to water increases its temperature until it boils. Every substance has its own characteristic ability to absorb heat. Some substances must absorb more heat than others to reach a certain temperature. These energy requirements for different substances are described in terms of a physical property called *specific heat*. **Specific heat** (*SH*) is the amount of heat needed to raise the temperature of exactly 1 g of a substance by exactly 1 °C. This temperature change is written as ΔT (*delta T*), where the delta symbol means "a change in."

$$\text{Specific heat } (SH) = \frac{\text{heat}}{\text{grams} \times \Delta T} = \frac{\text{J (or cal)}}{1\text{ g} \times 1\text{ °C}}$$

Now we can write the specific heat for water using our definition of the calorie and joule.

$$\text{Specific heat } (SH) \text{ of } H_2O(l) = \frac{4.184\text{ J}}{\text{g} \times \text{°C}} = \frac{1.00\text{ cal}}{\text{g} \times \text{°C}}$$

If we look at Table 2.7, we see that 1 g of water requires 4.184 J to increase its temperature by 1 °C. Water has a large specific heat that is about five times the specific heat of aluminum. Aluminum has a specific heat that is about twice that of copper.

TABLE 2.7 Specific Heats of Some Substances

Substance	J/g °C	cal/g °C
Elements		
Aluminum, Al(*s*)	0.897	0.214
Copper, Cu(*s*)	0.385	0.0920
Gold, Au(*s*)	0.129	0.0308
Iron, Fe(*s*)	0.452	0.108
Silver, Ag(*s*)	0.235	0.0562
Titanium, Ti(*s*)	0.523	0.125
Compounds		
Ammonia, $NH_3(g)$	2.04	0.488
Ethanol, $C_2H_5OH(l)$	2.46	0.588
Sodium chloride, NaCl(*s*)	0.864	0.207
Water, $H_2O(l)$	4.184	1.00
Water, $H_2O(s)$	2.03	0.485

Therefore, 4.184 J (or 1.00 cal) will increase the temperature of 1 g of water by 1 °C. However, adding the same amount of heat (4.184 J or 1.00 cal) will raise the temperature of 1 g of aluminum by about 5 °C and 1 g of copper by 10 °C. The low specific heats of aluminum and copper mean they transfer heat efficiently, which makes them useful in cookware.

The high specific heat of water has a major impact on the temperatures in a coastal city compared to an inland city. A large mass of water near a coastal city can absorb or release five times the energy absorbed or released by the same mass of rock near an inland city. This means that in the summer a body of water absorbs large quantities of heat, which cools a coastal city, and then in the winter that same body of water releases large quantities of heat, which provide warmer temperatures. A similar effect happens with our bodies, which contain 70% by mass water. Water in the body absorbs or releases large quantities of heat in order to maintain an almost constant body temperature.

CONCEPT CHECK 2.5

Comparing Specific Heats

Water has a specific heat that is about six times larger than that of sandstone. How would the temperature change during the day and night if you live in a house next to a large lake compared to a house built in the desert on sandstone?

ANSWER

In the day, the water in the lake will absorb six times more energy than sandstone, which will keep the temperature in a house on a large lake more comfortable and cooler than a house in the desert. In the night, the water in the lake will release energy that warms the surrounding air so that the temperature will not drop as much as in the desert.

TUTORIAL
Specific Heat Calculations

Calculations Using Specific Heat

When we know the specific heat of a substance, we can calculate the heat lost or gained by measuring the mass of the substance and the initial and final temperature. We can substitute these measurements into the specific heat expression that is rearranged to solve for heat, which we call the *heat equation.*

$$\begin{array}{rcccccc}
\text{Heat} & = & \text{mass} & \times & \text{temperature change} & \times & \text{specific heat} \\
\text{Heat} & = & \text{mass} & \times & \Delta T & \times & SH \\
\text{cal} & = & \text{grams} & \times & {}^\circ\text{C} & \times & \dfrac{\text{cal}}{\text{g}\ {}^\circ\text{C}} \\
\text{J} & = & \text{grams} & \times & {}^\circ\text{C} & \times & \dfrac{\text{J}}{\text{g}\ {}^\circ\text{C}}
\end{array}$$

SAMPLE PROBLEM 2.7

Calculating Heat with Temperature Increase

How many joules are absorbed by 45.2 g of aluminum if its temperature rises from 12.5 °C to 76.8 °C (see Table 2.7)?

SOLUTION

Step 1 **List given and needed data.**

Given mass = 45.2 g

SH for aluminum = 0.897 J/g °C

Initial temperature = 12.5 °C

Final temperature = 76.8 °C

Need heat in joules (J)

Step 2 **Calculate the temperature change (ΔT).** The temperature change, ΔT, is the difference between the two temperatures.

$$\Delta T = T_{\text{final}} - T_{\text{initial}} = 76.8\ °\text{C} - 12.5\ °\text{C} = 64.3\ °\text{C}$$

Step 3 **Write the heat equation and rearrange for unknown.**

$$\text{Heat} = \text{mass} \times \Delta T \times SH$$

Step 4 **Substitute the given values and solve, making sure units cancel.**

$$\text{Heat} = 45.2\ \cancel{\text{g}} \times 64.3\ \cancel{°\text{C}} \times \frac{0.897\ \text{J}}{\cancel{\text{g}}\ \cancel{°\text{C}}} = 2610\ \text{J}\ (2.61 \times 10^3\ \text{J})$$

STUDY CHECK 2.7

Some cooking pans have a layer of copper on the bottom. How many kilojoules are needed to raise the temperature of 125 g of copper from 22 °C to 325 °C (see Table 2.7)?

Guide to Calculations Using Specific Heat

1 List given and needed data.

2 Calculate the temperature change (ΔT).

3 Write the heat equation and rearrange for unknown.

4 Substitute the given values and solve, making sure units cancel.

The copper on a pan conducts heat rapidly to the food in the pan.

QUESTIONS AND PROBLEMS

Specific Heat

2.27 If the same amount of heat is supplied to samples of 10.0 g each of aluminum, iron, and copper all at 15 °C, which sample would reach the highest temperature (see Table 2.7)?

2.28 Substances **A** and **B** are the same mass and at the same initial temperature. When they are heated, the final temperature of **A** is 55 °C higher than the temperature of **B**. What does this tell you about the specific heats of **A** and **B**?

2.29 What is the amount of energy required in each of the following?
- **a.** calories to heat 8.5 g of water from 15 °C to 36 °C
- **b.** joules to heat 75 g of water from 22 °C to 66 °C
- **c.** kilocalories to heat 150 g of water in a kettle from 15 °C to 77 °C
- **d.** kilojoules to heat 175 g of copper from 28 °C to 188 °C

2.30 What is the amount of energy involved in each of the following?
- **a.** calories given off when 85 g of water cools from 45 °C to 25 °C
- **b.** joules given off when 25 g of water cools from 86 °C to 61 °C
- **c.** kilocalories added when 5.0 kg of water warms from 22 °C to 28 °C
- **d.** kilojoules to heat 224 g of gold from 18 °C to 185 °C

2.31 Calculate the energy, in joules and calories, for each of the following:
- **a.** required to heat 25.0 g of water from 12.5 °C to 25.7 °C
- **b.** required to heat 38.0 g of copper from 122 °C to 246 °C
- **c.** lost when 15.0 g of ethanol, C_2H_5OH, cools from 60.5 °C to −42.0 °C
- **d.** lost when 125 g of iron cools from 118 °C to 55 °C

2.32 Calculate the energy, in joules and calories, for each of the following:
- **a.** required to heat 5.25 g of water from 5.5 °C to 64.8 °C
- **b.** lost when 75.0 g of water cools from 86.4 °C to 2.1 °C
- **c.** required to heat 10.0 g of silver from 112 °C to 275 °C
- **d.** lost when 18.0 g of gold cools from 224 °C to 118 °C

LEARNING GOAL

Use the energy values to calculate the kilocalories (Cal) or kilojoules (kJ) in a food.

TUTORIAL
Nutritional Energy

CASE STUDY
Calories from Hidden Sugar

2.6 Energy and Nutrition

The food we eat provides energy to do work in the body, which includes the growth and repair of cells. Carbohydrates are the primary fuel for the body, but if the carbohydrate reserves are exhausted, fats, and then proteins are used for energy.

For many years in the field of nutrition, the energy from food was measured as Calories or kilocalories. The nutritional unit *Calorie*, *Cal* (with an uppercase C), is the same as 1000 cal, or 1 kcal. The international unit, kilojoule (kJ), is becoming more prevalent. For example, a baked potato has an energy value of 100 Calories, which is 100 kcal or 440 kJ. A typical diet that provides 2100 Cal (kcal) is the same as an 8800 kJ diet.

Energy Values in Nutrition

1 Cal = 1 kcal = 1000 cal

1 Cal = 4.184 kJ = 4184 J

The number of Calories in a food is determined by using an apparatus called a calorimeter, shown in Figure 2.7. A sample of food is placed in a steel container filled with oxygen with a measured amount of water which fills the surrounding chamber. The food sample is ignited, releasing heat that increases the temperature of the water. From the mass of the food and water as well as the temperature increase, the energy value of the food is calculated. We will assume that the energy absorbed by the calorimeter is negligible.

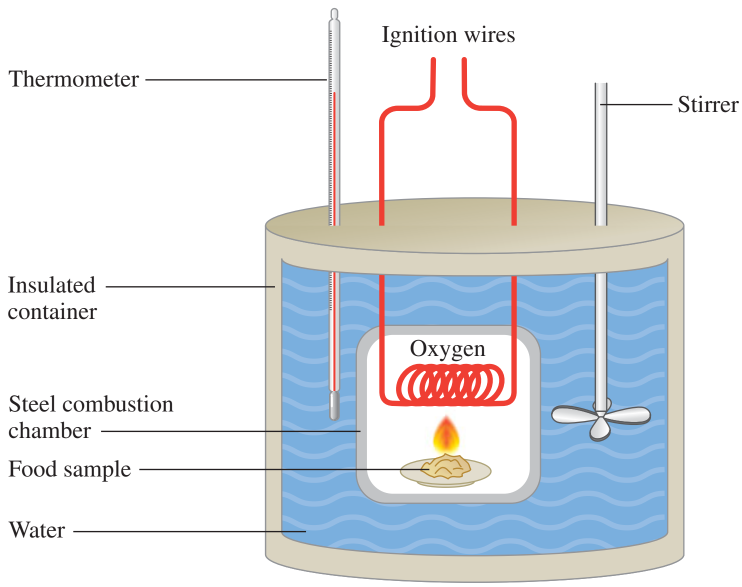


FIGURE 2.7 Heat released from burning a food sample in a calorimeter is used to determine the energy value of the food.

Q What happens to the temperature of water in a calorimeter during the combustion of a food sample?

CONCEPT CHECK 2.6

Energy Values of Food

When 55 g of pasta is burned in a calorimeter, 220 Cal of heat is released. What is the energy value of pasta in kcal/g?

ANSWER

Using the equality 1 Cal = 1 kcal, we can calculate the energy value of pasta.

$$\frac{220\,\cancel{\text{Cal}}}{55\text{ g}} \times \frac{1\text{ kcal}}{1\,\cancel{\text{Cal}}} = \frac{4.0\text{ kcal}}{\text{g}}$$

Energy Values for Foods

The **energy (caloric) values** of food are the kilocalories or kilojoules obtained from burning 1 g of carbohydrate, fat, or protein. These values are listed in Table 2.8.

Using the energy values in Table 2.8, we can calculate the total energy of a food if the mass of each food type is known.

TABLE 2.8 Typical Energy (Caloric) Values for the Three Food Types

Food Type	kJ/g	kcal/g
Carbohydrate	17	4
Fat	38	9
Protein	17	4

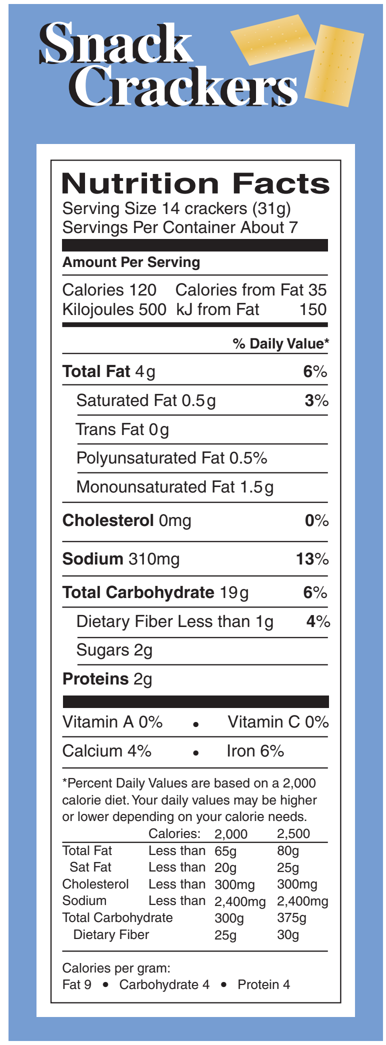
$$\text{Kilojoules} = \cancel{g} \times \frac{\text{kJ}}{\cancel{g}} \qquad \text{Kilocalories} = \cancel{g} \times \frac{\text{kcal}}{\cancel{g}}$$

On packaged food, the energy content is listed in the Nutrition Facts label, usually in terms of the number of Calories for one serving. The general composition and energy content of some foods are given in Table 2.9.

TABLE 2.9 General Composition and Energy Content of Some Foods

Food	Carbohydrate (g)	Fat (g)	Protein (g)	Energy*
Banana, 1 medium	26	0	1	460 kJ (110 kcal)
Beef, ground, 3 oz	0	14	22	910 kJ (220 kcal)
Carrots, raw, 1 cup	11	0	1	200 kJ (50 kcal)
Chicken, no skin, 3 oz	0	3	20	460 kJ (110 kcal)
Egg, 1 large	0	6	6	330 kJ (80 kcal)
Milk, 4% fat, 1 cup	12	9	9	700 kJ (170 kcal)
Milk, nonfat, 1 cup	12	0	9	360 kJ (90 kcal)
Potato, baked	23	0	3	440 kJ (100 kcal)
Salmon, 3 oz	0	5	16	460 kJ (110 kcal)
Steak, 3 oz	0	27	19	1350 kJ (320 kcal)

*Energy values are rounded off to the tens place.



The nutrition facts include the total Calories, Calories from fat, and total grams of carbohydrate.

SAMPLE PROBLEM 2.8

Caloric Content for a Food

At a fast-food restaurant, a hamburger contains 37 g of carbohydrate, 19 g of fat, and 24 g of protein. What is the total energy content in kilocalories? Round off the kilocalories for each type of food to the tens place.

SOLUTION

Using the energy values for carbohydrate, fat, and protein (see Table 2.8), we can calculate the kilocalories for each type of food and the total kcal:

Food Type	Mass		Energy Value		Energy
Carbohydrate	$37\ \cancel{g}$	$\times$	$\frac{4\text{ kcal}}{1\ \cancel{g}}$	$=$	150 kcal*
Fat	$19\ \cancel{g}$	$\times$	$\frac{9\text{ kcal}}{1\ \cancel{g}}$	$=$	170 kcal*
Protein	$24\ \cancel{g}$	$\times$	$\frac{4\text{ kcal}}{1\ \cancel{g}}$	$=$	100 kcal*
	Total energy content			$=$	420 kcal

*Energy results are rounded off to the tens place.

STUDY CHECK 2.8

If you buy the same hamburger as in Sample Problem 2.8 at a fast-food restaurant in Canada, what is the energy content stated in kilojoules? Round off the kilojoules for each food type to the tens place.

Explore Your World

COUNTING CALORIES

Obtain a food item that has a nutrition label. From the information on the label, determine the number of grams of carbohydrate, fat, and protein in one serving. Using the caloric values, calculate the total Calories for one serving. (For most products, the kilocalories for each food type are rounded off to the tens place.)

QUESTION

How does your total for the Calories in one serving compare to the Calories stated on the label for a single serving?

Chemistry Link to Health

LOSING AND GAINING WEIGHT

The number of kilocalories needed in the daily diet of an adult depends on gender and physical activity. Some general levels of energy needs are given in Table 2.10.

The amount of food a person eats is regulated by the hunger center in the hypothalamus, which is located in the brain. Food intake is normally proportional to the nutrient stores in the body. If these nutrient stores are low, you feel hungry; if they are high, you do not feel like eating.

A person gains weight when food intake exceeds energy output and loses weight when food intake is less than energy output. Many diet products contain cellulose, which has no nutritive value but provides bulk and makes you feel full. Some diet drugs depress the hunger center and must be used with caution, because they excite the nervous system and can elevate blood pressure. Because muscular exercise is an important way to expend energy, an increase in daily exercise aids weight loss. Table 2.11 lists some activities and the amount of energy they require.

One hour of swimming uses 2100 kJ of energy.

TABLE 2.10 Typical Energy Requirements for Adults

Gender	Age	Moderately active kcal (kJ)	Highly active kcal (kJ)
Female	19–30	2100 (8800)	2400 (10 000)
	31–50	2000 (8400)	2200 (9200)
Male	19–30	2700 (11 300)	3000 (12 600)
	31–50	2500 (10 500)	2900 (12 100)

TABLE 2.11 Energy Expended by a 70.0-kg (154-lb) Adult

Activity	Energy (kcal/h)	Energy (kJ/h)
Sleeping	60	250
Sitting	100	420
Walking	200	840
Swimming	500	2100
Running	750	3100

QUESTIONS AND PROBLEMS

Energy and Nutrition

2.33 Using the following data, calculate the kilocalories for each food burned in a calorimeter:
a. one stalk of celery that heats 505 g of water from 25.2 °C to 35.7 °C
b. a waffle that heats 4980 g of water from 20.6 °C to 62.4 °C

2.34 Using the following data, calculate the kilocalories for each food burned in a calorimeter:
a. 1 cup of popcorn that heats 1250 g of water from 25.5 °C to 50.8 °C
b. a sample of butter that heats 357 g of water from 22.7 °C to 38.8 °C

2.35 Using the energy values for foods (see Table 2.8), determine each of the following (round off the answers in kilocalories or kilojoules to the tens place):
a. the kilojoules for 1 cup of orange juice that has 26 g of carbohydrate, 2 g of protein, and no fat
b. the grams of carbohydrate in one apple if the apple has no fat and no protein and provides 72 Cal of energy
c. the kilocalories in 1 tablespoon of vegetable oil, which contains 14 g of fat and no carbohydrate or protein
d. the total kilocalories for a diet that consists of 68 g of carbohydrate, 150 g of protein, and 9.0 g of fat

2.36 Using the energy values for foods (see Table 2.8), determine each of the following (round off the answers in kilocalories or kilojoules to the tens place):
a. the kilojoules in 2 tablespoons of crunchy peanut butter that contains 6 g of carbohydrate, 16 g of fat, and 7 g of protein
b. the grams of protein in a cup of soup that has 110 kcal with 7 g of fat and 9 g of carbohydrate
c. the grams of sugar (carbohydrate) in a can of cola that has 140 Cal and no fat and no protein
d. the grams of fat in one avocado that has 405 kcal, 13 g of carbohydrate, and 5 g of protein

2.37 One cup of clam chowder contains 9 g of protein, 12 g of fat, and 16 g of carbohydrate. How much energy, in kilocalories and kilojoules, is in the clam chowder? (Round off the kilocalories or kilojoules to the tens place.)

2.38 A high-protein diet contains 70. g of carbohydrate, 150 g of protein, and 5.0 g of fat. How many kilocalories does this diet provide? How many kilojoules does this diet provide? (Round off the kilocalories and kilojoules to the tens place.)

2.7 Changes of State

LEARNING GOAL

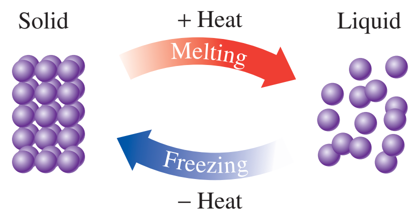
Describe the changes of state between solids, liquids, and gases; calculate the energy involved.

In Section 2.2, we described the properties and states of matter: gases, liquids, and solids. We can now discuss how matter undergoes a **change of state** when it is converted from one state to another (see Figure 2.8).

When heat is added to a solid, the particles move faster. At a temperature called the **melting point (mp)**, the particles of a solid gain sufficient energy to overcome the attractive forces that hold them together. The particles in the solid separate and move about in random patterns. The substance is **melting**, changing from a solid to a liquid.

If the temperature is lowered, the reverse process takes place. Kinetic energy is lost, the particles slow down, and attractive forces pull the particles close together. The substance is **freezing**. A liquid changes to a solid at the **freezing point (fp)**, which is the same temperature as its melting point. Every substance has its own freezing (melting) point: water freezes (melts) at 0 °C; gold freezes (melts) at 1064 °C; nitrogen freezes (melts) at −210 °C.

During a change of state, the temperature of a substance remains constant. Suppose we have a glass containing ice and water. The ice melts when heat is added at 0 °C, forming more liquid. The liquid freezes when heat is removed at 0 °C. The processes of melting and freezing are reversible at 0 °C.



Melting and freezing are reversible processes.

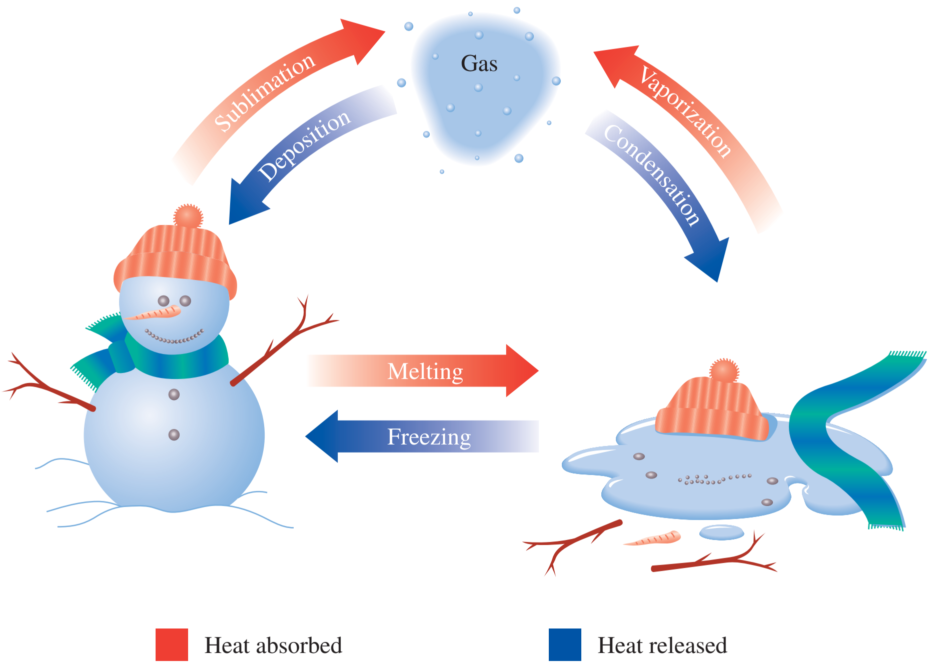


FIGURE 2.8 A summary of the changes of state.

Q Is heat added or released when liquid water freezes?

Heat of Fusion

During melting, energy called the **heat of fusion** is needed to separate the particles of a solid. For example, 80. cal (334 J) of heat are needed to melt exactly 1 g of ice at its melting point (0 °C).

Heat of Fusion for Water

$$\frac{80.\ \text{cal}}{1\ \text{g water}} \qquad \frac{334\ \text{J}}{1\ \text{g water}}$$

The heat of fusion (80. cal/g or 334 J/g) is also the heat that must be removed to freeze exactly 1 g of water at its freezing point (0 °C). Water is sometimes sprayed in fruit orchards during subfreezing weather. If the air temperature drops to 0 °C, the water begins to freeze. Heat is released as the water freezes, which warms the air and protects the fruit.

To determine the heat needed to melt a sample of ice, multiply the mass of the ice by its heat of fusion. There is no temperature change in the calculation because temperature remains constant as long as the ice is melting.

Calculating Heat to Melt (or Freeze) Water

Heat = mass × heat of fusion

$$\text{cal} = \cancel{\text{g}} \times \frac{80.\ \text{cal}}{\cancel{\text{g}}} \qquad \text{J} = \cancel{\text{g}} \times \frac{334\ \text{J}}{\cancel{\text{g}}}$$

SAMPLE PROBLEM 2.9

Heat of Fusion

Ice cubes at 0 °C with a mass of 26 g are added to your soft drink.

a. How much heat (cal) must be added to melt all the ice at 0 °C?
b. What happens to the temperature of your soft drink? Why?

Guide to Calculations Using Heat of Fusion

1 List the grams of substance and change of state.

2 Write the plan to convert grams to heat and desired unit.

3 Write the heat conversion factor and metric factor if needed.

4 Set up the problem with factors.

SOLUTION

a. The heat in calories required to melt the ice is calculated as follows:

Step 1 List the grams of substance and change of state.

Given 26 g of $H_2O(s)$ **Need** calories to melt ice

Step 2 Write the plan to convert grams to heat and desired unit.

grams of ice → Heat of fusion → calories

Step 3 Write the heat conversion factor and metric factor if needed.

$$1\ \text{g of } H_2O(s \rightarrow l) = 80.\ \text{cal}$$

$$\frac{80.\ \text{cal}}{1\ \text{g } H_2O} \quad \text{and} \quad \frac{1\ \text{g } H_2O}{80.\ \text{cal}}$$

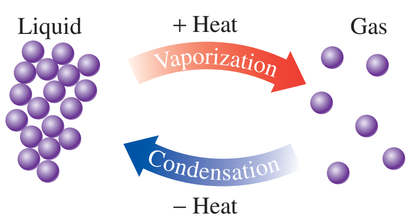
Step 4 Set up the problem with factors.

$$26\ \cancel{\text{g } H_2O} \times \frac{80.\ \text{cal}}{1\ \cancel{\text{g } H_2O}} = 2100\ \text{cal}$$

b. The soft drink will be colder because heat from the soft drink is providing the energy to melt the ice.

STUDY CHECK 2.9

In a freezer, 150 g of water at 0 °C is placed in an ice cube tray. How much energy, in kilocalories, must be removed to form ice cubes at 0 °C?



Vaporization and condensation are reversible processes.

Boiling and Condensation

Water in a mud puddle disappears, unwrapped food dries out, and clothes hung on a line dry. **Evaporation** is taking place as water molecules with sufficient energy escape from the liquid surface and enter the gas phase (see Figure 2.9a). The loss of the "hot" water molecules removes heat, which cools the remaining liquid water. As heat is added, more and more water molecules evaporate. At the **boiling point (bp)**, the molecules of the liquid have the energy needed to change into a gas. The **boiling** of a liquid occurs as gas bubbles form throughout the liquid, then rise to the surface and escape (see Figure 2.9b).

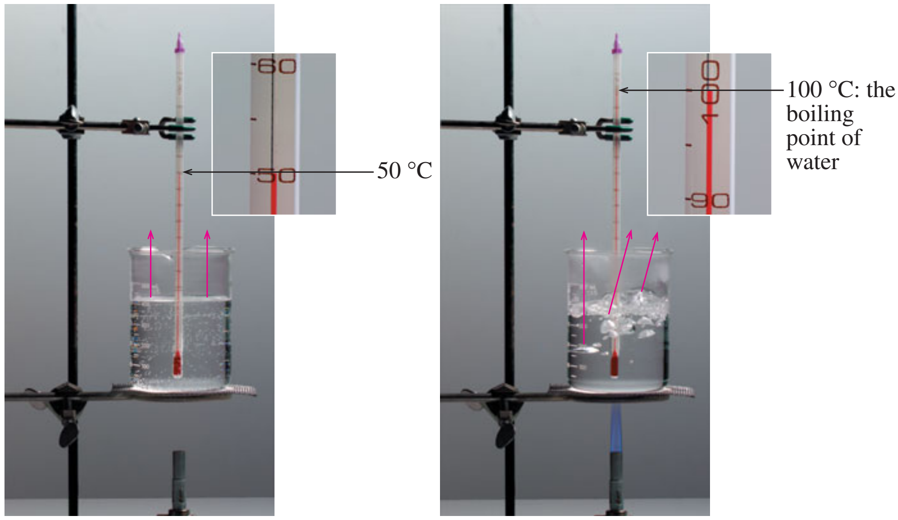


FIGURE 2.9 (a) Evaporation occurs at the surface of a liquid. (b) Boiling occurs as bubbles of gas form throughout the liquid.
Q Why does water evaporate faster at 80 °C than at 20 °C?

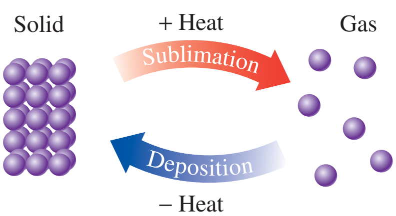
When heat is removed, a reverse process takes place. In **condensation**, water vapor is converted back to liquid as the water molecules lose kinetic energy and slow down. Condensation occurs at the same temperature as boiling but differs because heat is removed. You may have noticed that condensation occurs when you take a hot shower and the water vapor forms water droplets on a mirror. Because a substance loses heat as it condenses, its surroundings become warmer. That is why, when a rainstorm is approaching, you may notice a warming of the air as gaseous water molecules condense to rain.

Sublimation

In a process called **sublimation**, the particles on the surface of a solid change directly to a gas without going through the liquid state. In the process called **deposition**, gas changes directly to a solid. For example, dry ice, which is solid carbon dioxide, sublimes at −78 °C. It is called "dry" because it does not form a liquid as it warms. In extremely cold areas, snow does not melt but sublimes directly to water vapor.

Heat of Sublimation for Water

$$\frac{620.\ \text{cal}}{1\ \text{g water}} \qquad \frac{2590\ \text{J}}{1\ \text{g water}}$$



Sublimation and deposition are reversible processes.

Dry ice sublimes at −78 °C.

Career Focus

HISTOLOGIST

"While a patient is in surgery for skin cancer, some tissue around the cancer is sent to us," says Mary Ann Pipe, histology technician. "Using the Mohs surgical technique, we place the tissue block on a glass slide, chill it to −30 °C in a machine called a cryostat, and freeze it for longer in another machine called a heat extractor. From this frozen block of tissue, we cut extremely thin slices—one-thousandth of an inch—from different depths. We prepare three separate slides from skin at three different depths up to the surface of the skin. We stain the cells pink and blue by placing the slides in hemotoxin, and then in eosin. The slices are a tissue map that the doctor can easily read to determine if the margins around the skin cancer are clear or whether more tissue must be removed."

Water vapor will change to solid on contact with a cold surface.

Freeze-dried foods have a long shelf life because they contain no water.

TUTORIAL
Heat of Vaporization and Heat of Fusion

When frozen foods are left in the freezer for a long time, so much water sublimes that foods, especially meats, become dry and shrunken, a condition called *freezer burn*. Deposition occurs in a freezer when water vapor forms ice crystals on the surface of freezer bags and frozen food.

Freeze-dried foods prepared by sublimation are convenient for long-term storage and for camping and hiking. A food that has been frozen is placed in a vacuum chamber where it dries as the ice sublimes. The dried food retains all of its nutritional value and needs only water to be edible. A food that is freeze-dried does not need refrigeration because bacteria cannot grow without moisture.

CONCEPT CHECK 2.7

Identifying Changes of State

Give the change of state described in the following:

a. particles on the surface of a liquid escaping to form vapor
b. a liquid changing to a solid
c. gas bubbles forming throughout a liquid

ANSWER

a. evaporation **b.** freezing **c.** boiling

Heat of Vaporization

The energy that must be added to convert exactly 1 g of liquid to gas at its boiling point is called the **heat of vaporization**. For water, 540 cal (2260 J) is needed to convert 1 g of water to vapor at 100 °C. This same amount of heat is released when 1 g of water vapor (gas) changes to liquid at 100 °C. Therefore, 2260 J or 540 cal/g is also the *heat of condensation* of water.

Heat of Vaporization for Water

$$\frac{540\text{ cal}}{1\text{ g water}} \qquad \frac{2260\text{ J}}{1\text{ g water}}$$

To calculate the amount of heat added to (or removed from) a sample of water, the mass of the sample is multiplied by the heat of vaporization. As before, no temperature change occurs during this change of state.

Calculating Heat to Vaporize (or Condense) Water

Heat = mass × heat of vaporization

$$\text{cal} = \cancel{g} \times \frac{540\text{ cal}}{\cancel{g}} \qquad \text{J} = \cancel{g} \times \frac{2260\text{ J}}{\cancel{g}}$$

Guide to Calculations Using Heat of Vaporization

1. List the grams of substance and change of state.
2. Write the plan to convert grams to heat and desired unit.
3. Write the heat conversion factor and metric factor if needed.
4. Set up the problem with factors.

SAMPLE PROBLEM 2.10

Using Heat of Vaporization

In a sauna, 122 g of water is converted to steam at 100 °C. How many kilojoules of heat are needed?

SOLUTION

Step 1 List the grams of substance and change of state.

Given 122 g of $H_2O(l)$ to $H_2O(g)$

Need kilojoules of heat to change state

Step 2 **Write the plan to convert grams to heat and desired unit.**

grams of water $\xrightarrow{\text{Heat of vaporization}}$ calories $\xrightarrow{\text{Metric factor}}$ kilocalories

Step 3 **Write the heat conversion factor and metric factor if needed.**

$$1 \text{ g of } H_2O\ (l \rightarrow g) = 2260 \text{ J}$$
$$\frac{2260 \text{ J}}{1 \text{ g } H_2O} \quad \text{and} \quad \frac{1 \text{ g } H_2O}{2260 \text{ J}}$$

$$1 \text{ kJ} = 1000 \text{ J}$$
$$\frac{1000 \text{ J}}{1 \text{ kJ}} \quad \text{and} \quad \frac{1 \text{ kJ}}{1000 \text{ J}}$$

Step 4 **Set up the problem with factors.**

$$122\ \cancel{\text{g } H_2O} \times \frac{2260\ \cancel{\text{J}}}{1\ \cancel{\text{g } H_2O}} \times \frac{1 \text{ kJ}}{1000\ \cancel{\text{J}}} = 276 \text{ kJ}$$

STUDY CHECK 2.10

When steam from a pan of boiling water reaches a cool window, it condenses. How much heat, in kilojoules, is released when 25.0 g of steam condenses at 100 °C?

Chemistry Link to Health

STEAM BURNS

Hot water at 100 °C will cause burns and damage to the skin. However, getting steam on the skin is even more dangerous. If 25 g of hot water at 100 °C falls on a person's skin, the temperature of the water will drop to body temperature, 37 °C. The heat released during cooling can cause severe burns. The amount of heat can be calculated from the temperature change, 100 °C − 37 °C = 63 °C.

$$25\ \cancel{\text{g}} \times 63\ \cancel{°\text{C}} \times \frac{4.184 \text{ J}}{\cancel{\text{g}}\ \cancel{°\text{C}}} = 6600 \text{ J}$$

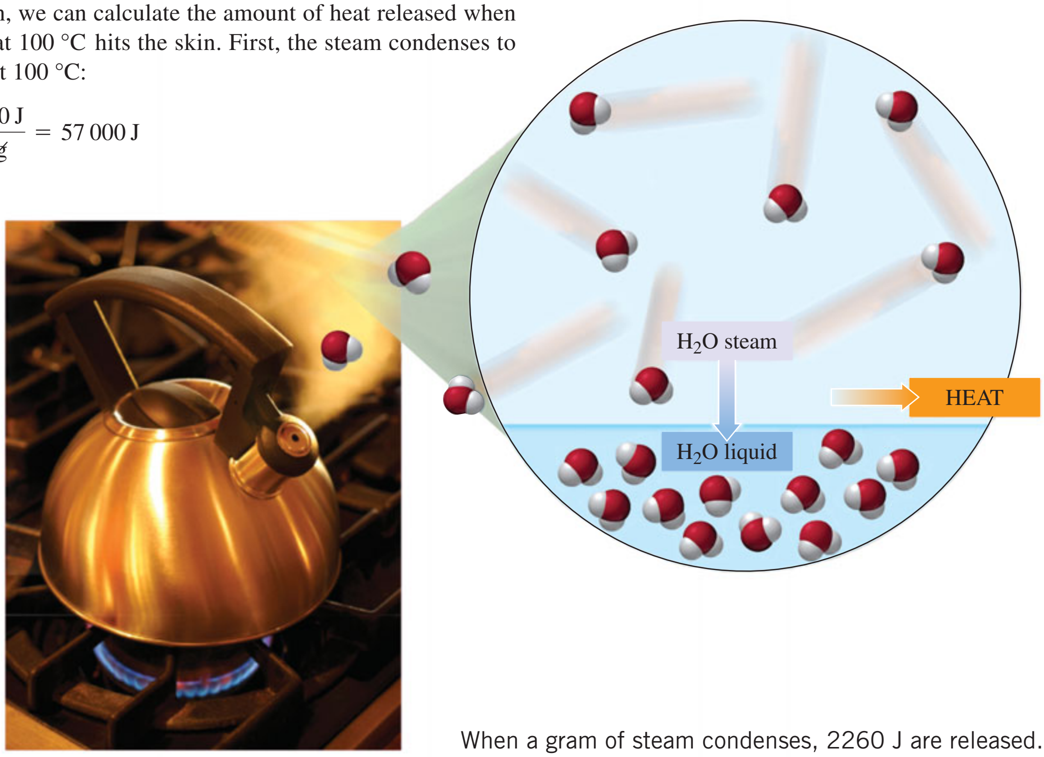
For comparison, we can calculate the amount of heat released when 25 g of steam at 100 °C hits the skin. First, the steam condenses to water (liquid) at 100 °C:

$$25\ \cancel{\text{g}} \times \frac{2260 \text{ J}}{1\ \cancel{\text{g}}} = 57\,000 \text{ J}$$

Then, the liquid water at 100 °C cools to body temperature 37 °C. The total amount of heat released from the condensation and cooling of the steam is calculated as follows:

Condensation (100 °C)	= 57 000 J
Cooling (100 °C to 37 °C)	= 6 600 J
Heat released	= 64 000 J (rounded off)

The amount of heat released from steam is almost ten times greater than the heat from the same amount of hot water.



When a gram of steam condenses, 2260 J are released.

Heating and Cooling Curves

All the changes of state during the heating of a solid can be illustrated visually. On a **heating curve**, the temperature is shown on the vertical axis and the addition of heat is shown on the horizontal axis (see Figure 2.10a).

Steps on a Heating Curve

The first diagonal line indicates a warming of a solid as heat is added. When the melting temperature is reached, a horizontal line, or plateau, indicates that the solid is melting. As melting takes place, the solid is changing to liquid without any change in temperature (see Figure 2.10a).

Once all the particles are in the liquid state, adding more heat will increase the temperature of the liquid. This increase is shown as a diagonal line from the melting point temperature to the boiling point temperature. Once the liquid reaches its boiling point, a horizontal line indicates that the temperature is constant as liquid changes to gas. Because the heat of vaporization is greater than the heat of fusion, the horizontal line at the boiling point is longer than the line at the melting point. Once all the liquid becomes gas, adding more heat increases the temperature of the gas.

Steps on a Cooling Curve

A **cooling curve** is a diagram of the cooling process in which the temperature decreases as heat is removed (see Figure 2.10b). The cooling of the gas is shown as a diagonal line to the boiling (condensation) point. At the boiling (condensation) point, the horizontal line indicates a change of state as gas condenses to form a liquid. When all the gas has changed into liquid, the temperature decreases with further cooling of the liquid, which is shown as a diagonal line from the condensation point temperature to the freezing point temperature. At the freezing point, another horizontal line indicates that liquid is changing to solid at the freezing point temperature. Once all the substance is frozen, the removal of more heat lowers the temperature below its freezing point, which is shown as a diagonal line below the freezing point.

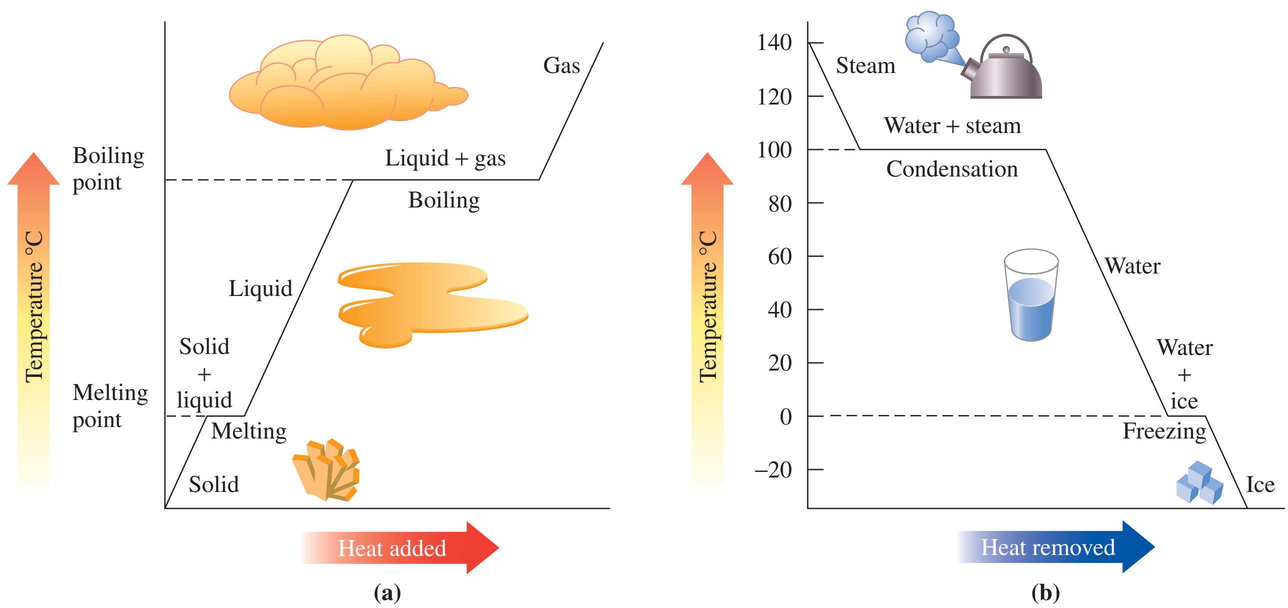


FIGURE 2.10 (a) A heating curve diagrams changes in state as temperature increases. (b) A cooling curve for water diagrams changes in state as temperature decreases.

Q What does the plateau at 100 °C represent on the cooling curve for water?

CONCEPT CHECK 2.8

Using a Cooling Curve

Using the cooling curve for water in Figure 2.10b, identify the state or change of state for water as solid, liquid, gas, condensation, or freezing.

a. at 120 °C **b.** at 100 °C
c. at 40 °C **d.** at 0 °C

ANSWER

a. A temperature of 120 °C occurs on the diagonal line above the boiling (condensation) point indicating that water is a gas.
b. A temperature of 100 °C, shown as a horizontal line, indicates that the water vapor is changing to liquid water, or condensing.
c. A temperature of 40 °C occurs on the diagonal line below the boiling point but above the freezing point, which indicates that the water is in the liquid state.
d. A temperature of 0 °C, shown as a horizontal line, indicates that the liquid water is changing to solid (ice) or freezing.

Combining Energy Calculations

Up to now, we have calculated one step in a heating or cooling curve. However, many problems require a combination of steps that include a temperature change as well as a change of state. The heat is calculated for each step separately and then added together to find the total energy, as seen in Sample Problem 2.11.

MC

TUTORIAL
Heat, Energy, and Changes of State

SAMPLE PROBLEM 2.11

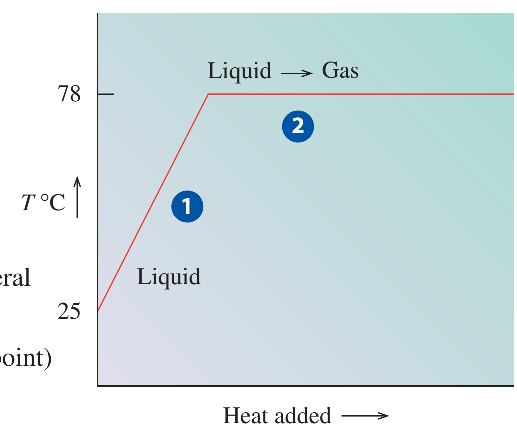
Combining Heat Calculations

Calculate the total heat, in joules, needed to convert 15.0 g of liquid ethanol at 25.0 °C to gas at its boiling point of 78.0 °C. Ethanol has a specific heat of 2.46 J/g °C, and a heat of vaporization of 841 J/g.

SOLUTION

Step 1 List the grams of substance and change of state.

Given 15.0 g of ethanol at 25.0 °C; boiling point of ethanol 78.0 °C
Specific heat 2.46 J/g °C; heat of vaporization 841 J/g
Need heat (J) to warm ethanol and change to gas at the boiling point



Step 2 Write the plan to convert grams to heat and desired unit. When several changes occur, draw a diagram of heating and changes of state.

Total heat = joules needed to warm ethanol from 25.0 °C to 78.0 °C (boiling point)
+ joules to change liquid to gas at 78.0 °C (boiling point)

Step 3 Write the heat conversion factors needed.

$$SH_{\text{Ethanol}} = \frac{2.46\ \text{J}}{\text{g}\ °\text{C}}$$

$$\frac{2.46\ \text{J}}{\text{g}\ °\text{C}} \quad \text{and} \quad \frac{\text{g}\ °\text{C}}{2.46\ \text{J}}$$

$$1\ \text{g of ethanol}\ (l \longrightarrow g) = 841\ \text{J}$$

$$\frac{841\ \text{J}}{1\ \text{g ethanol}} \quad \text{and} \quad \frac{1\ \text{g ethanol}}{841\ \text{J}}$$

Step 4 Set up problem with factors.

$\Delta T = 78.0\ °C - 25.0\ °C = 53.0\ °C$

Heat needed to warm ethanol (liquid) at 25 °C to ethanol (liquid) at 78.0 °C (boiling point):

$$15.0\ \cancel{g} \times 53.0\ \cancel{°C} \times \frac{2.46\ J}{\cancel{g}\ \cancel{°C}} = 1960\ J$$

Heat needed to change ethanol (liquid) to ethanol (gas) at 78.0 °C (boiling point):

$$15.0\ \cancel{\text{g ethanol}} \times \frac{841\ J}{1\ \cancel{\text{g ethanol}}} = 12\ 600\ J$$

Calculate the total heat:

Heating ethanol (25.0 °C to boiling point 78.0 °C)	1 960 J
Changing liquid to gas at boiling point (78.0 °C)	12 600 J
Total heat needed	14 600 J (rounded off)

STUDY CHECK 2.11

How many kilojoules are released when 75.0 g of steam at 100 °C condenses, cools to 0 °C, and freezes at 0 °C? (*Hint:* The solution will require three energy calculations.)

QUESTIONS AND PROBLEMS

Changes of State

2.39 Identify each of the following changes of state as melting, freezing, sublimation, or deposition:
a. The solid structure of a substance breaks down as liquid forms.
b. Coffee is freeze-dried.
c. Water on the street turns to ice during a cold wintry night.
d. Ice crystals form on a package of frozen corn.

2.40 Identify each of the following changes of state as melting, freezing, sublimation, or deposition:
a. Dry ice in an ice-cream cart disappears.
b. Snow on the ground turns to liquid water.
c. Heat is removed from 125 g of liquid water at 0 °C.
d. Frost (ice) forms on the walls of a freezer unit of a refrigerator.

2.41 Calculate the heat needed at 0 °C in each of the following and indicate whether heat was absorbed or released:
a. calories to melt 65 g of ice
b. joules to melt 17.0 g of ice
c. kilocalories to freeze 225 g of water
d. kilojoules to freeze 50.0 g of water

2.42 Calculate the heat needed at 0 °C in each of the following and indicate whether heat was absorbed or released:
a. calories to freeze 35 g of water
b. joules to freeze 275 g of water
c. kilocalories to melt 140 g of ice
d. kilojoules to melt 5.00 kg of ice

2.43 Identify each of the following changes of state as evaporation, boiling, or condensation:
a. The water vapor in the clouds changes to rain.
b. Wet clothes dry on a clothesline.
c. Lava flows into the ocean and steam forms.
d. After a hot shower, your bathroom mirror is covered with water.

2.44 Identify each of the following changes of state as evaporation, boiling, or condensation:
a. At 100 °C, the water in a pan changes to steam.
b. On a cool morning, the windows in your car fog up.
c. A shallow pond dries up in the summer.
d. Your teakettle whistles when the water is ready for tea.

2.45 Calculate the heat change at 100 °C in each of the following problems. Indicate whether heat was absorbed or released.
a. calories to vaporize 10.0 g of water
b. joules to vaporize 5.00 g of water
c. kilocalories to condense 8.0 kg of steam
d. kilojoules to condense 175 g of steam

2.46 Calculate the heat change at 100 °C in each of the following problems. Indicate whether heat was absorbed or released.
a. calories to condense 10.0 g of steam
b. joules to condense 7.60 g of steam
c. kilocalories to vaporize 44 g of water
d. kilojoules to vaporize 5.00 kg of water

2.47 Draw a heating curve for a sample of ice that is heated from −20 °C to 150 °C. Indicate the segment of the graph that corresponds to each of the following:
a. solid **b.** melting **c.** liquid
d. boiling **e.** gas

2.48 Draw a cooling curve for a sample of steam that cools from 110 °C to −10 °C. Indicate the segment of the graph that corresponds to each of the following:
a. solid
b. freezing
c. liquid
d. condensation (boiling)
e. gas

2.49 Using the values for the heat of fusion, specific heat of water, and/or heat of vaporization, calculate the amount of heat energy in each of the following:

a. joules needed to melt 50.0 g of ice at 0 °C and to warm the liquid to 65.0 °C

b. kilocalories released when 15.0 g of steam condenses at 100 °C and the liquid cools to 0 °C

c. kilojoules needed to melt 24.0 g of ice at 0 °C, warm the liquid to 100 °C and change it to steam at 100 °C

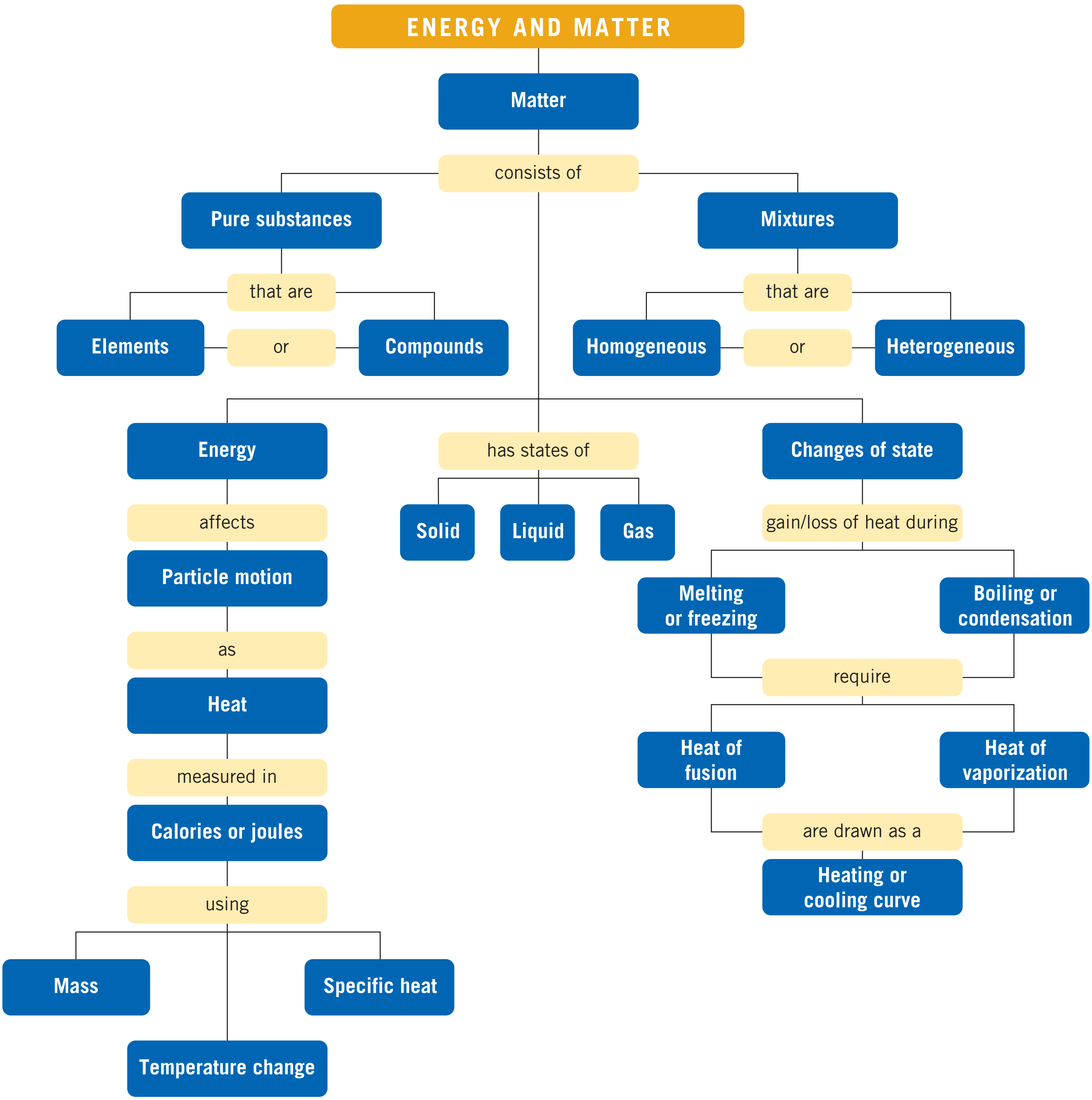
2.50 Using the values for the heat of fusion, specific heat of water, and/or heat of vaporization, calculate the amount of heat energy in each of the following:

a. joules to condense 125 g of steam at 100 °C and to cool the liquid to 15.0 °C

b. kilocalories needed to melt a 525-g ice sculpture at 0 °C and to warm the liquid to 15.0 °C

c. kilojoules released when 85.0 g of steam condenses at 100 °C and to cool the liquid and freeze it at 0 °C

CONCEPT MAP



CHAPTER REVIEW

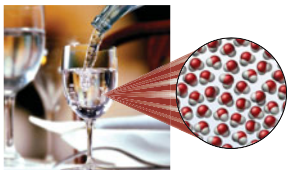
2.1 Classification of Matter

Learning Goal: Classify examples of matter as pure substances or mixtures.

Matter is classified as pure substances or mixtures. Pure substances, which are elements or compounds, have fixed compositions. Mixtures have variable compositions, which are classified further as homogeneous or heterogeneous. The substances in mixtures can be separated using physical methods.

2.2 States and Properties of Matter

Learning Goal: Identify the states and physical and chemical properties of matter.



The three states of matter are solid, liquid, and gas. A physical property is a characteristic that can be observed without changing the identity of the substance. A physical change occurs when physical properties but not the composition of the substance, change. A chemical property describes the ability of a substance to change into a different substance. In a chemical change, at least one substance forms a new substance with new physical properties.

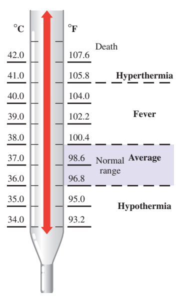
2.3 Energy

Learning Goal: Identify energy as potential or kinetic; convert between units of energy.

Energy is the ability to do work. Kinetic energy is the energy of motion; potential energy is the energy determined by position or composition. Common units of energy are the calorie (cal), kilocalorie (kcal), joule (J), and kilojoule (kJ).

2.4 Temperature

Learning Goal: Given a temperature, calculate a corresponding temperature on another scale.



On the Celsius scale, there are 100 units between the freezing point of water (0 °C) and the boiling point (100 °C). On the Fahrenheit scale, there are 180 units between the freezing point of water (32 °F) and the boiling point (212 °F). A Fahrenheit temperature is related to its Celsius temperature by the equation $T_F = 1.8\,T_C + 32$. The SI temperature of Kelvin is related to the Celsius temperature by the equation $T_K = T_C + 273$.

2.5 Specific Heat

Learning Goal: Use specific heat to calculate heat loss or gain, temperature change, or mass of a sample.

Specific heat is the amount of energy required to raise the temperature of exactly 1 g of a substance by exactly 1 °C. The heat lost or gained by a substance is determined by multiplying its mass, the temperature change, and its specific heat.

2.6 Energy and Nutrition

Learning Goal: Use the energy values to calculate the kilocalories (Cal) or kilojoules (kJ) in a food.

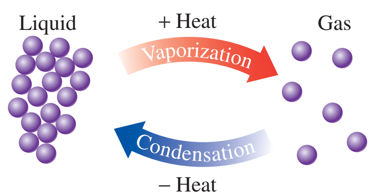
The nutritional Calorie (Cal) is the same amount of energy as 1 kcal or 1000 calories. The energy content of a food is the sum of kilocalories or kilojoules from carbohydrate, fat, and protein.

TABLE 2.8 Typical Energy (Caloric) Values for the Three Food Types

Food Type	kJ/g	kcal/g
Carbohydrate	17	4
Fat	38	9
Protein	17	4

2.7 Changes of State

Learning Goal: Describe the changes of state between solids, liquids, and gases; calculate the energy involved.



Melting occurs when the particles in a solid absorb enough energy to break apart and form a liquid. The amount of energy required to convert exactly 1 g of solid to liquid is called the heat of fusion. For water, 80. cal (334 J) are needed to melt 1 g of ice. Boiling is the vaporization of a liquid at its boiling point. The heat of vaporization is the amount of heat needed to convert exactly 1 g of liquid to vapor. For water, 540 cal (2260 J) are needed to vaporize 1 g of liquid water. A heating or cooling curve illustrates the changes in temperature and state as heat is added to or removed from a substance. Plateaus on the graph indicate changes of state with no change in temperature.

Key Terms

boiling The formation of bubbles of gas throughout a liquid.

boiling point (bp) The temperature at which a liquid changes to gas (boils) and gas changes to liquid (condenses).

calorie (cal) The amount of heat energy that raises the temperature of exactly 1 g of water exactly 1 °C.

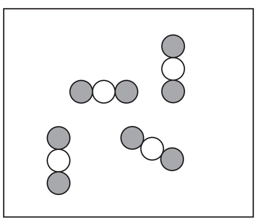
change of state The transformation of one state of matter to another; for example, solid to liquid, liquid to solid, liquid to gas.

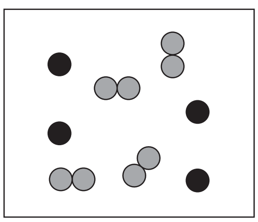
chemical change A change during which the original substance is converted into a new substance that has a different composition and new chemical and physical properties.

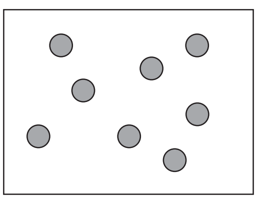
chemical properties The properties that indicate the ability of a substance to change into a new substance.
compound A pure substance consisting of two or more elements, with a definite composition, that can be broken down into simpler substances only by chemical methods.
condensation The change of state of a gas to a liquid.
cooling curve A diagram that illustrates temperature changes and changes of state for a substance as heat is removed.
deposition The change of a gas directly into a solid; the reverse of sublimation.
element A pure substance containing only one type of matter, which cannot be broken down by chemical methods.
energy The ability to do work.
energy (caloric) value The kilocalories (or kilojoules) obtained per gram of the food types: carbohydrate, fat, and protein.
evaporation The formation of a gas (vapor) by the escape of high-energy molecules from the surface of a liquid.
freezing The change of state from liquid to solid.
freezing point (fp) The temperature at which a liquid changes to a solid (freezes), a solid changes to a liquid (melts).
gas A state of matter that does not have a definite shape or volume.
heat The energy associated with the motion of particles in a substance.
heat of fusion The energy required to melt exactly 1 g of a substance at its melting point. For water, 80. cal (334 J) are needed to melt 1 g of ice; 80. cal (334 J) are released when 1 g of water freezes.
heat of vaporization The energy required to vaporize exactly 1 g of a substance at its boiling point. For water, 540 calories (2260 J) are needed to vaporize 1 g of liquid; 1 g of steam gives off 540 cal (2260 J) when it condenses.
heating curve A diagram that shows the temperature changes and changes of state of a substance as it is heated.
joule (J) The SI unit of heat energy; 4.184 J = 1 cal.
kinetic energy The energy of moving particles.
liquid A state of matter that takes the shape of its container but has a definite volume.
matter The material that makes up a substance and has mass and occupies space.
melting The change of state from a solid to a liquid.
melting point (mp) The temperature at which a solid becomes a liquid (melts). It is the same temperature as the freezing point.
mixture The physical combination of two or more substances that does not change the identities of the mixed substances.
physical change A change in which the physical properties of a substance change but its identity stays the same.
physical properties The properties that can be observed or measured without affecting the identity of a substance.
potential energy A type of energy related to position or composition of a substance.
pure substance A type of matter composed of elements and compounds that has a definite composition.
solid A state of matter that has its own shape and volume.
specific heat A quantity of heat that changes the temperature of exactly 1 g of a substance by exactly 1 °C.
states of matter Three forms of matter: solid, liquid, and gas.
sublimation The change of state in which a solid is transformed directly to a gas without forming a liquid first.

Understanding the Concepts

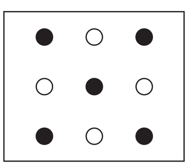
2.51 Identify each of the following as an element, a compound, or a mixture:

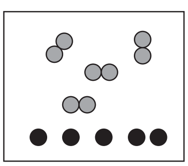
a.


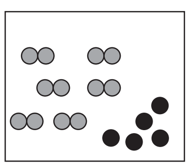
b.


c.


2.52 Which diagram illustrates a homogeneous mixture? Explain your choice. Which of the following diagrams illustrates heterogeneous mixtures? Explain your choice.


a.

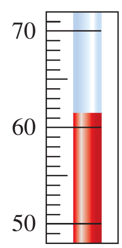

b.


c.

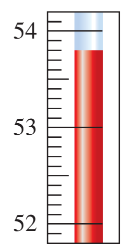
2.53 Classify each of the following as a homogeneous or heterogeneous mixture:
a. lemon-flavored water **b.** stuffed mushrooms
c. tortilla soup

2.54 Classify each of the following as a homogeneous or heterogeneous mixture:
a. ketchup **b.** hard-boiled egg
c. eye drops

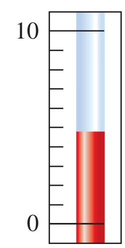
2.55 State the temperature, including the estimated digit, on each of the Celsius thermometers.



A



B



C

2.56 Select the warmer temperature in each pair.

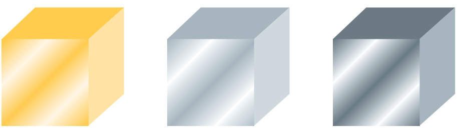
a. 10 °C or 10 °F
b. 30 °C or 15 °F
c. −10 °C or 32 °F
d. 200 °C or 200 K

2.57 Compost can be made at home from grass clippings, some kitchen scraps, and dry leaves. As microbes break down organic matter, heat is generated and the compost can reach a temperature of 155 °F, which kills most pathogens. What is this initial temperature in Celsius degrees? In kelvins?

Compost produced from decayed plant material is used to enrich the soil.

2.58 After a week, biochemical reactions in compost slow, and the temperature drops to 45 °C. The dark brown organic-rich mixture is ready for use in the garden. What is this temperature in Fahrenheit degrees? In kelvins?

2.59 Calculate the energy to heat three cubes (gold, aluminum, and silver) each with a volume of $10.0\ \text{cm}^3$ from 15 °C to 25 °C. Refer to Tables 1.13 and 2.7. What do you notice about the energy needed for each?



2.60 If you used the 8400 kilojoules you expend in energy in one day to heat 50 000 g of water at 20 °C, what would be the rise in temperature? What would be the new temperature of the water?

2.61 A 70.0-kg person has just eaten a quarter-pound cheeseburger, french fries, and a chocolate shake. Using Table 2.8, calculate the total kilocalories in this meal (round off the kilocalories to the tens place).

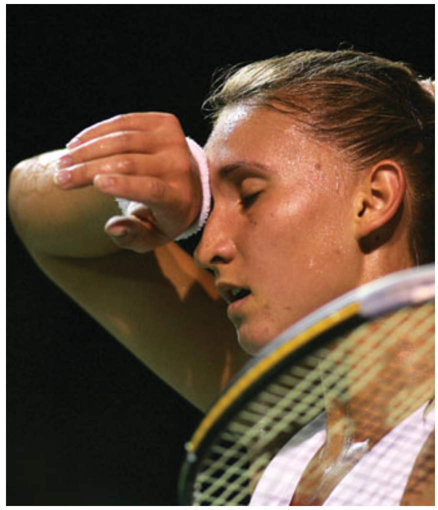
Item	Protein (g)	Fat (g)	Carbohydrate (g)
Cheeseburger	31	29	34
French fries	3	11	26
Chocolate shake	11	9	60

2.62 For the person in Problem 2.61, use Table 2.11 to determine each of the following:

a. The number of hours of sleeping needed to "burn off" the kilocalories in this meal.
b. The number of hours of running needed to "burn off" the kilocalories in this meal.

2.63 Use your knowledge of changes of state to explain the following:

a. How does perspiration during heavy exercise cool the body?
b. Why do towels dry more quickly on a hot summer day than on a cold winter day?
c. Why do wet clothes stay wet in a plastic bag?



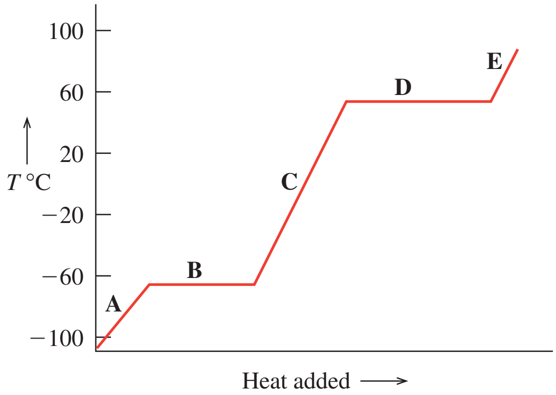
Perspiration forms on the skin during heavy exercise.

2.64 Use your knowledge of changes of state to explain the following:

a. Why is a spray, such as ethyl chloride, used to numb a sports injury during a game?
b. Why does water in a wide, flat, shallow dish evaporate more quickly than the same amount of water in a tall, narrow glass?
c. Why does a sandwich on a plate dry out faster than a sandwich in plastic wrap?

A spray is used to numb a sports injury.

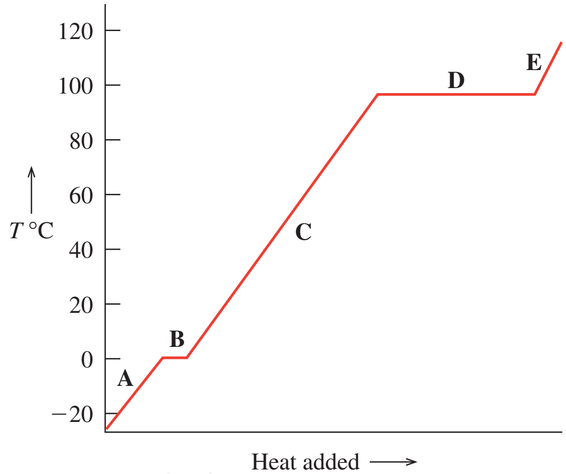
2.65 The following graph is a heating curve for chloroform, a solvent for fats, oils, and waxes:

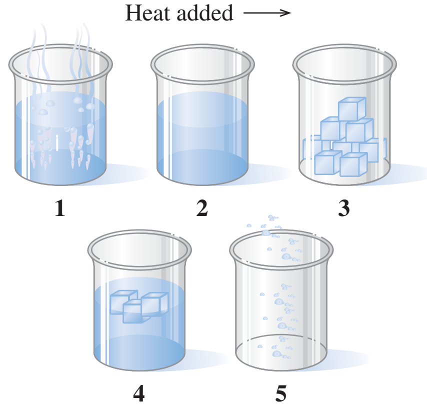


a. What is the approximate melting point of chloroform?
b. What is the approximate boiling point of chloroform?
c. On the heating curve, identify the segments **A**, **B**, **C**, **D**, and **E** as solid, liquid, gas, melting, or boiling.
d. At the following temperatures, is chloroform a solid, liquid, or gas?

−80 °C; −40 °C; 25 °C; 80 °C

2.66 Associate the contents of the beakers (**1–5**) with segments (**A–E**) on the following heating curve for water:





Additional Questions and Problems

For instructor-assigned homework, go to **www.masteringchemistry.com**.

2.67 Classify each of the following as an element, a compound, or a mixture:
a. carbon in pencils
b. carbon dioxide (CO_2) we exhale
c. orange juice
d. neon gas in lights
e. salad dressing of oil and vinegar

2.68 Classify each of the following as a homogeneous or heterogeneous mixture:
a. hot fudge sundae
b. herbal tea
c. vegetable oil
d. water and sand
e. mustard

2.69 Identify each of the following as a solid, a liquid, or a gas:
a. vitamin tablets in a bottle
b. helium in a balloon
c. milk in a glass
d. the air you breathe
e. charcoal briquettes on a barbecue

2.70 Identify each of the following as a solid, a liquid, or a gas:
a. popcorn in a bag
b. water in a garden hose
c. a computer mouse
d. air in a tire
e. hot tea

2.71 Identify each of the following as a physical or chemical property:
a. Gold is shiny.
b. Gold melts at 1064 °C.
c. Gold is a good conductor of electricity.
d. When gold reacts with yellow sulfur, a black sulfide compound forms.

2.72 Identify each of the following as a physical or chemical property of a candle:
a. The candle is 20 cm high with a diameter of 3 cm.
b. The candle burns.
c. The wax of the candle softens on a hot day.
d. The candle is blue.

2.73 Identify each of the following as a physical or chemical change:
a. A plant grows a new leaf.
b. Chocolate is melted for a dessert.
c. Wood is chopped for the fireplace.
d. Wood burns in a fireplace.

2.74 Identify each of the following as a physical or chemical change:
a. A medication tablet is broken in two.
b. Carrots are grated for use in a salad.
c. Malt undergoes fermentation to make beer.
d. A copper pipe reacts with air and turns green.

2.75 Calculate each of the following temperatures in degrees Celsius:
a. The highest recorded temperature in the continental United States was 134 °F in Death Valley, California, on July 10, 1913.
b. The lowest recorded temperature in the continental United States was −69.7 °F in Rodgers Pass, Montana, January 20, 1954.

2.76 Calculate each the following temperatures in degrees Fahrenheit:
a. The highest recorded temperature in the world was 58.0 °C in El Azizia, Libya, on September 13, 1922.
b. The lowest recorded temperature in the world was −89.2 °C in Vostok, Antarctica, on July 21, 1983.

2.77 What is −15 °F in degrees Celsius and in kelvins?

2.78 The highest recorded body temperature that a person has survived is 46.5 °C. Calculate that temperature in degrees Fahrenheit and in kelvins.

2.79 On a hot day, the beach sand gets hot, but the water stays cool. Would you predict the specific heat of sand is higher or lower than that of water? Explain.

The water, sand, and air gain energy from the Sun.

2.80 Why do drops of liquid water form on the outside of a glass of iced tea?

2.81 If you want to lose 1 pound of "fat," which is 15% water, how many kilocalories do you need to lose?

2.82 Calculate the Cal (kcal) in 1/2 cup of soft ice cream that contains 18 g of carbohydrate, 11 g of fat, and 4 g of protein. (Round off the kilocalories to the tens place.)

2.83 A hot-water bottle contains 725 g of water at 65 °C. If the water cools to body temperature (37 °C), how many kilocalories of heat could be transferred to sore muscles?

2.84 A pitcher containing 0.75 L of water at 4 °C is removed from the refrigerator. How many kilojoules are needed to warm the water to a room temperature of 22 °C?

2.85 The melting point of chloroform is −64 °C and its boiling point is 61 °C. Sketch a heating curve for chloroform from −100 °C to 100 °C.
a. What is the state of chloroform at −75 °C?
b. What happens on the curve at −64 °C?
c. What is the state of chloroform at −18 °C?
d. What is the state of chloroform at 80 °C?
e. At what temperature will both solid and liquid be present?

2.86 The melting point of benzene is 5.5 °C and its boiling point is 80.1 °C. Sketch a heating curve for benzene from 0 °C to 100 °C.
a. What is the state of benzene at 15 °C?
b. What happens on the curve at 5.5 °C?
c. What is the state of benzene at 63 °C?
d. What is the state of benzene at 98 °C?
e. At what temperature will both liquid and gas be present?

Challenge Questions

2.87 One liquid has a temperature of 140. °F and another liquid has a temperature of 60.0 °C. Are the liquids at the same temperature or at different temperatures?

2.88 A 0.50-g sample of vegetable oil is placed in a calorimeter. When the sample is burned, 18.9 kJ are given off. What is the caloric value, in kcal/g, of the oil?

2.89 How many kilocalories of heat are released when 75 g of steam at 100 °C is converted to ice at 0 °C? (*Hint*: The calculations include several steps.)

2.90 A 45-g piece of ice at 0.0 °C is added to a sample of water at 8.0 °C. All of the ice melts and the temperature of the water decreases to 0.0 °C. How many grams of water were in the sample?

2.91 In a large building, oil is used in a steam boiler heating system. The combustion of 1.0 lb of oil provides 2.4×10^7 J.
a. How many kilograms of oil are needed to heat 150 kg of water from 22 °C to 100 °C?
b. How many kilograms of oil are needed to change 150 kg of water to steam at 100 °C?

2.92 When 1.0 g of gasoline burns, it releases 11 kcal of heat. The density of gasoline is 0.74 g/mL.
a. How many megajoules are released when 1.0 gal of gasoline burns?
b. If a television requires 150 kJ/h to run, how many hours can the television run on the energy provided by 1.0 gal of gasoline?

2.93 An ice bag containing 275 g of ice at 0 °C was used to treat sore muscles. When the bag was removed, the ice had melted and the liquid water had a temperature of 24.0 °C. How many kilojoules of heat were absorbed?

2.94 A 115-g sample of steam at 100 °C is emitted from a volcano. It condenses, cools, and falls as snow at 0 °C. How many kilojoules of heat were released?

Answers

Answers to Study Checks

2.1 heterogeneous

2.2 **b** and **d** are chemical changes.

2.3 35 000 J

2.4 12 °F

2.5 39.8 °C

2.6 night −260. °C; day 410. °C

2.7 14.6 kJ

2.8 1760 kJ

2.9 12 kcal

2.10 56.5 kJ released

2.11 226 kJ

Answers to Selected Questions and Problems

2.1 **a.** pure substance **b.** mixture **c.** pure substance **d.** pure substance **e.** mixture

2.3 **a.** element **b.** compound **c.** element **d.** compound **e.** compound

2.5 **a.** heterogeneous **b.** homogeneous **c.** homogeneous **d.** heterogeneous **e.** heterogeneous

2.7 **a.** gas **b.** gas **c.** solid

2.9 **a.** physical **b.** chemical **c.** physical **d.** chemical **e.** chemical

2.11 **a.** physical **b.** chemical **c.** physical **d.** physical **e.** physical

2.13 **a.** chemical **b.** physical **c.** physical **d.** chemical **e.** physical

2.15 When the roller-coaster car is at the top of the ramp, it has its maximum potential energy. As it descends, potential energy changes to kinetic energy. At the bottom, all the energy is kinetic.

2.17 **a.** potential **b.** kinetic **c.** potential **d.** potential

2.19 **a.** 8.1×10^5 J **b.** 190 kcal

2.21 In the United States, we still use the Fahrenheit temperature scale. In °F, normal body temperature is 98.6. On the Celsius scale, her temperature would be 37.7 °C, a mild fever.

2.23 **a.** 98.6 °F **b.** 18.5 °C **c.** 246 K **d.** 335 K **e.** 46 °C

2.25 **a.** 41 °C **b.** No. The temperature is equivalent to 39 °C.

2.27 Copper has the lowest specific heat of the samples and will reach the highest temperature.

2.29 **a.** 180 cal **b.** 14 000 J **c.** 9.3 kcal **d.** 10.8 kJ

2.31 **a.** 1380 J, 330. cal **b.** 1810 J, 434 cal **c.** 3780 J, 904 cal **d.** 3600 J, 850 cal

2.33 **a.** 5.30 kcal **b.** 208 kcal

2.35 **a.** 470 kJ **b.** 18 g **c.** 130 kcal **d.** 950 kcal

2.37 210 kcal, 880 kJ

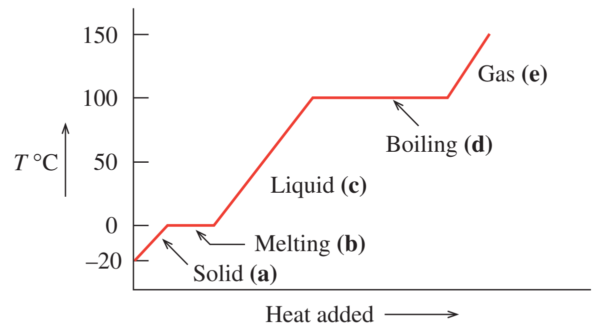
2.39 **a.** melting **b.** sublimation **c.** freezing **d.** deposition

2.41 **a.** 5200 cal absorbed
b. 5680 J absorbed
c. 18 kcal released
d. 16.7 kJ released

2.43 **a.** condensation **b.** evaporation **c.** boiling **d.** condensation

2.45 **a.** 5400 cal absorbed
b. 11 300 J absorbed
c. 4300 kcal released
d. 396 kJ released

2.47



2.49 **a.** 30 300 J **b.** 9.6 kcal **c.** 72.2 kJ

2.51 **a.** compound **b.** mixture **c.** element

2.53 **a.** homogeneous **b.** heterogeneous **c.** heterogeneous

2.55 **a.** 61.4 °C **b.** 53.80 °C **c.** 4.8 °C

2.57 68.3 °C, 341 K

2.59 gold, 250 J or 59 cal; aluminum, 240 J or 58 cal; silver, 250 J or 59 cal. The heat needed for all the metal samples is almost the same.

2.61 1100 kcal (rounded off)

2.63 **a.** The heat from the skin is used to evaporate the water (perspiration). Therefore, the skin is cooled.
b. On a hot day, there are more molecules with sufficient energy to become water vapor.
c. In a closed plastic bag, some water molecules evaporate, but they cannot escape and will condense back to liquid; the clothes will not dry.

2.65 **a.** about −60 °C
b. about 60 °C
c. The diagonal line A represents the solid state as temperature increases. The horizontal line B represents the change from

solid to liquid or melting of the substance. The diagonal line C represents the liquid state as temperature increases. The horizontal line D represents the change from liquid to gas or boiling of the liquid. The diagonal line E represents the gas state as temperature increases.

d. At −80 °C, solid; at −40 °C, liquid; at 25 °C, liquid; at 80 °C, gas

2.67 **a.** element **b.** compound **c.** mixture **d.** element **e.** mixture

2.69 **a.** solid **b.** gas **c.** liquid **d.** gas **e.** solid

2.71 **a.** physical **b.** physical **c.** physical **d.** chemical

2.73 **a.** chemical **b.** physical **c.** physical **d.** chemical

2.75 **a.** 57 °C (or 56.7 °C using 3 SFs) **b.** −56.5 °C

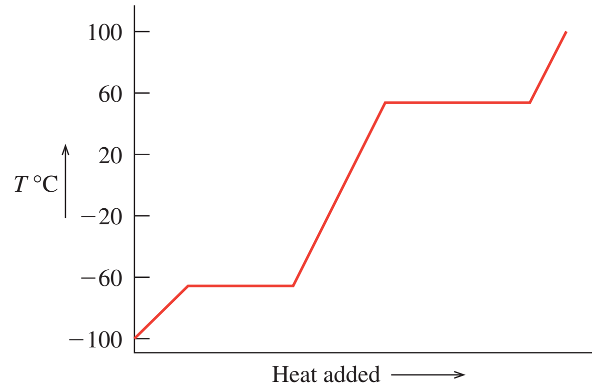
2.77 −26 °C; 247 K

2.79 Sand must have a lower specific heat than water. When both substances absorb the same amount of heat, the final temperature of the sand will be higher than that of water.

2.81 3500 kcal

2.83 20. kcal

2.85



a. solid **b.** solid chloroform melts **c.** liquid **d.** gas **e.** −64 °C

2.87 The two liquids are at the same temperature.

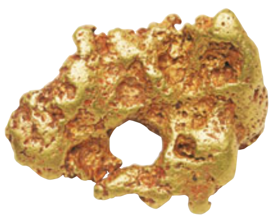
2.89 55 kcal

2.91 **a.** 0.93 kg **b.** 6.4 kg

2.93 119.5 kJ

Combining Ideas from Chapters 1 and 2

CI.1 Gold, one of the most sought-after metals in the world, has a density of 19.3 g/cm^3, a melting point of 1064 °C, a specific heat of 0.129 J/g °C, and a heat of fusion of 63.6 J/g. A gold nugget found in Alaska in 1998 weighs 20.17 lb.



Gold nuggets, also called native gold, can be found in streams and mines.

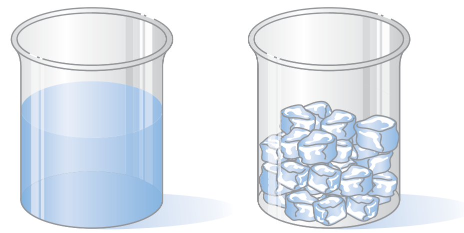
a. How many significant figures are in the measurement of the weight of the nugget?
b. Which is the mass of the nugget in kilograms?
c. If the nugget were pure gold, what would its volume be in cm^3?
d. What is the melting point of gold in Fahrenheit degrees and kelvins?
e. How many kilocalories are required to raise the temperature of the nugget from 500. °C to 1064 °C and melt all the gold to liquid at 1064 °C?
f. If the price of gold is \$35.10 per gram, what is the nugget worth, in dollars?

CI.2 The mileage for a motorcycle with a fuel-tank capacity of 22 L is 35 mi/gal. The density of gasoline is 0.74 g/mL.

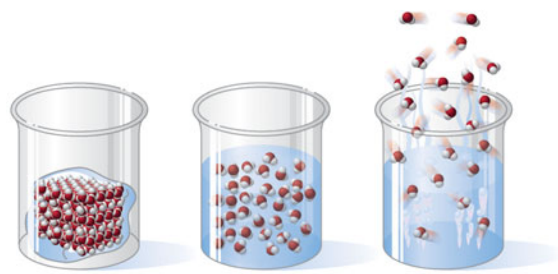
a. How long a trip, in kilometers, can be made on one full tank of gasoline?
b. If the price of gasoline is \$2.67 per gallon, what would be the cost of fuel for the trip?
c. If the average speed during the trip is 44 mi/h, how many hours will it take to reach the destination?
d. If the density of gasoline is 0.74 g/mL, what is the mass, in grams, of the fuel in the tank?
e. When 1.00 g of gasoline burns, 47 kJ of energy is released. How many kilojoules are produced when the fuel in one full tank is burned?

CI.3 Answer the following questions for the water samples A and B shown in the diagrams:

A **B**



a. When each sample is transferred to another container, what happens to the shape and volume of each?
b. Match the diagrams (1, 2, or 3) that represent the water particles with sample **A** or **B**. Give a reason for your choice.



1 2 3

c. The state of matter indicated in diagram 1 is a __________; in diagram 2 it is a __________; and in diagram 3 it is a __________.
d. When the water in diagram 1 changes to the water in diagram 2, the process is called __________, which occurs at a temperature called the __________ point. This is an example of a __________ change.
e. When the water in diagram 2 changes to the water in diagram 3, the process is called __________, which occurs at a temperature called the __________ point. This is an example of a __________ change.
f. If the water in diagram 2 has a mass of 19 g and a temperature of 45 °C, how much heat, in kilojoules, is removed to cool the liquid and form solid at 0 °C?

CI.4 The label of a black cherry almond energy bar with a mass of 68 g lists the "nutrition facts" as 5 g of fat, 39 g of carbohydrate, and 10 g of protein.

a. Using the energy values of carbohydrates, fats, and proteins (see Table 2.8), what are the total kilocalories (Calories) listed for a black cherry almond bar? (Round off answers for each food type to the tens place.)
b. What are the kilojoules for the black cherry almond bar? (Round off answers for each food type to the tens place.)
c. If you obtain 160 kJ, how many grams of the black cherry almond bar did you eat?
d. If you are walking and using energy at a rate of 840 kJ/h, how many minutes of walking will you need to walk to expend the energy of two bars?

CI.5 In a box of nails, there are 75 iron nails weighing 0.25 lb. The density of iron is 7.86 g/cm^3. The specific heat of iron is 0.452 J/g °C. The melting point of iron is 1535 °C. The heat of fusion for iron is 272 J/g.

a. What is the volume, in cm^3, of the iron nails in the box?
b. If 30 nails are added to a graduated cylinder containing 17.6 mL of water, what is the new level of water in the cylinder?
c. How many joules must be added to the nails in the box to raise the temperature from 16 °C to 125 °C?
d. How many joules are required to heat one nail from 25 °C to its melting point and change it to liquid iron?

CI.6 A hot tub is filled with 450 gal of water.

a. What is the volume of water, in liters, in the tub?
b. What is the mass, in kilograms, of water in the tub?
c. How many kilocalories are needed to heat the water from 62 °F to 105 °F?
d. If the hot-tub heater provides 5900 kJ/min, how long, in hours, will it take to heat the water in the hot tub from 62 °F to 105 °F?

Answers

CI.1 **a.** 4 significant figures
b. 9.17 kg
c. 474 cm^3
d. 1947 °F; 1337 K
e. 298 kcal
f. $322 000

CI.3 **a.** The shape of A changes to the shape of the new container, while the shape of B remains the same. The volumes of both A and B remain the same.
b. A is liquid water represented by diagram 2. In liquid water, the water particles are in a random arrangement, but close together. B is solid water represented by diagram 1. In solid water, the water particles are fixed in a definite arrangement.
c. solid; liquid; gas
d. melting; melting point; physical
e. boiling; boiling point; physical
f. 9.9 kJ

CI.5 **a.** 14 cm^3
b. 23.4 mL
c. 5600 J or 5.6×10^3 J
d. 1400 J

Atoms and Elements

3

LOOKING AHEAD

"Many of my patients have diabetes, ulcers, hypertension, and cardiovascular problems," says Sylvia Lau, registered dietitian. "If a patient has diabetes, I discuss foods that raise blood sugar such as fruit, milk, and starches. I talk about how dietary fat contributes to weight gain and complications from diabetes. For stroke patients, I suggest diets low in fat and cholesterol because high blood pressure increases the risk of another stroke."

If a lab test shows low levels of iron, zinc, iodine, magnesium, or calcium, a dietitian discusses foods that provide those essential elements. For instance, she may recommend more beef for an iron deficiency, whole grain for zinc, leafy green vegetables for magnesium, dairy products for calcium, and iodized table salt and seafood for iodine.

All matter is composed of *elements*, of which there are 117 different kinds. Of these, 88 elements occur naturally and make up all the substances in our world. Many elements are already familiar to you. You may have a ring or necklace made of gold, silver, or perhaps platinum. If you play tennis or golf, then you may have noticed that your racket or clubs may be made from the elements titanium or carbon. In our bodies, calcium and phosphorus form the structure of bones and teeth, iron and copper are needed in the formation of red blood cells, and iodine is required for the proper functioning of the thyroid.

The correct amounts of certain elements are crucial to the proper growth and function of the body. Low levels of iron can lead to anemia, while lack of iodine can cause hypothyroidism and goiter. Some elements known as microminerals, such as chromium, cobalt, and selenium, are needed in our bodies in very small amounts. Laboratory tests are used to confirm that these elements are within normal ranges in our bodies.

LEARNING GOAL

Given the name of an element, write its correct symbol; from the symbol, write the correct name.

MC TUTORIAL
Elements and Symbols in the Periodic Table

3.1 Elements and Symbols

Elements are pure substances from which all other things are built. As we discussed in Chapter 2, elements cannot be broken down into simpler substances. Over the centuries, elements have been named for planets, mythological figures, minerals, colors, geographic locations, and famous people. Some sources of names of elements are listed in Table 3.1. A complete list of all the elements and their symbols appears on the inside front cover of this text.

TABLE 3.1 Some Elements and Their Names

Element	Source of Name
Uranium	The planet Uranus
Titanium	Titans (mythology)
Chlorine	*Chloros*, "greenish-yellow" (Greek)
Iodine	*Ioeides*, "violet" (Greek)
Magnesium	Magnesia, a mineral
Californium	California
Curium	Marie and Pierre Curie
Copernicium	Nicolaus Copernicus

Chemical Symbols

Chemical symbols are one- or two-letter abbreviations for the names of the elements. Only the first letter of an element's symbol is capitalized. If the symbol has a second letter, it is lowercase so that we know when a different element is indicated. If two letters are capitalized, they represent the symbols of two different elements. For example, the element cobalt has the symbol Co. However, the two capital letters CO specify two elements, carbon (C) and oxygen (O).

One-Letter Symbols	Two-Letter Symbols
C carbon	Co cobalt
S sulfur	Si silicon
N nitrogen	Ne neon
I iodine	Ni nickel

Although most of the symbols use letters from the current names, some are derived from their ancient names. For example, Na, the symbol for sodium, comes from the Latin word *natrium*. The symbol for iron, Fe, is derived from the Latin name *ferrum*. Table 3.2 lists the names and symbols of some common elements. Learning their names and symbols will greatly help your learning of chemistry.

TABLE 3.2 Names and Symbols of Some Common Elements

Name*	Symbol	Name*	Symbol	Name*	Symbol
Aluminum	Al	Gold (*aurum*)	Au	Oxygen	O
Argon	Ar	Helium	He	Phosphorus	P
Arsenic	As	Hydrogen	H	Platinum	Pt
Barium	Ba	Iodine	I	Potassium (*kalium*)	K
Boron	B	Iron (*ferrum*)	Fe	Radium	Ra
Bromine	Br	Lead (*plumbum*)	Pb	Silicon	Si
Cadmium	Cd	Lithium	Li	Silver (*argentum*)	Ag
Calcium	Ca	Magnesium	Mg	Sodium (*natrium*)	Na
Carbon	C	Manganese	Mn	Strontium	Sr
Chlorine	Cl	Mercury (*hydrargyrum*)	Hg	Sulfur	S
Chromium	Cr	Neon	Ne	Tin (*stannum*)	Sn
Cobalt	Co	Nickel	Ni	Titanium	Ti
Copper (*cuprum*)	Cu	Nitrogen	N	Uranium	U
Fluorine	F			Zinc	Zn

*Names given in parentheses are ancient Latin or Greek words from which the symbols are derived.


Aluminum


Carbon


Gold


Silver


Sulfur

CONCEPT CHECK 3.1

Symbols of the Elements

The symbol for carbon is C, and the symbol for sulfur is S. However, the symbol for cesium is Cs, not CS. Why?

ANSWER

When the symbol for an element has two letters, the first letter is capitalized, but the second letter is lowercase. If both letters are capitalized such as in CS, two elements—carbon and sulfur—are indicated.

SAMPLE PROBLEM 3.1

Writing Chemical Symbols

What are the chemical symbols for the following elements?

a. nickel **b.** nitrogen **c.** neon

SOLUTION

a. Ni **b.** N **c.** Ne

STUDY CHECK 3.1

What are the chemical symbols for the elements silicon, sulfur, and silver?

Chemistry Link to Health

TOXICITY OF MERCURY

Mercury is a silvery, shiny element that is a liquid at room temperature. Mercury can enter the body through inhaled mercury vapor, contact with the skin, or ingestion of foods or water contaminated with mercury. In the body, mercury destroys proteins and disrupts cell function. Long-term exposure to mercury can damage the brain and kidneys, cause mental retardation, and decrease physical development. Blood, urine, and hair samples are used to test for mercury.

In both freshwater and seawater, bacteria convert mercury into toxic methylmercury, which attacks the central nervous system (CNS). Because fish absorb methylmercury, we are exposed to mercury by consuming mercury-contaminated fish. As levels of mercury ingested from fish became a concern, the Food and Drug Administration (FDA) set a maximum level of one part mercury per million parts seafood (1 ppm), which is the same as 1 mg of mercury in every kilogram of seafood. Fish higher in the food chain, such as swordfish and shark, can have such high levels of mercury that the Environmental Protection Agency (EPA) recommends they be consumed no more than once a week.

One of the worst incidents of mercury poisoning occurred in Minamata and Niigata, Japan, in 1950. At that time, the ocean was polluted with high levels of mercury from industrial wastes. Because fish were a major food in the diet, more than 2000 people were affected with mercury poisoning and died or developed neural damage. In the United States between 1988 and 1997, the use of mercury decreased by 75% when the use of mercury was banned in paint and pesticides, and regulated in batteries and other products. Mercury batteries come with warnings on the label and should be disposed of properly.

This mercury fountain, housed in glass, was designed by Alexander Calder for the 1937 World's Fair in Paris.

SAMPLE PROBLEM 3.2

Names and Symbols of Chemical Elements

Give the name of the element that corresponds to each of the following chemical symbols:

a. Zn **b.** K **c.** H **d.** Fe

SOLUTION

a. zinc **b.** potassium **c.** hydrogen **d.** iron

STUDY CHECK 3.2

What are the names of the elements with the chemical symbols Mg, Al, and F?

QUESTIONS AND PROBLEMS

Elements and Symbols

3.1 Write the symbols for the following elements:
a. copper **b.** platinum **c.** calcium **d.** manganese
e. iron **f.** barium **g.** lead **h.** strontium

3.2 Write the symbols for the following elements:
a. oxygen **b.** lithium **c.** uranium **d.** titanium
e. hydrogen **f.** chromium **g.** tin **h.** gold

3.3 Write the name of the element for each symbol.
a. C **b.** Cl **c.** I **d.** Hg
e. Ag **f.** Ar **g.** B **h.** Ni

3.4 Write the name of the element for each symbol.
a. He **b.** P **c.** Na **d.** As
e. Ca **f.** Br **g.** Cd **h.** Si

3.5 What elements are in the following substances?
a. table salt, NaCl **b.** plaster casts, $CaSO_4$
c. Demerol, $C_{15}H_{22}ClNO_2$ **d.** antacid, $CaCO_3$

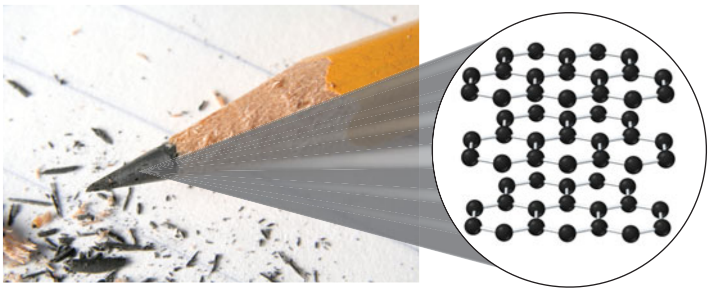
3.6 What elements are in the following substances?
a. water, H_2O **b.** baking soda, $NaHCO_3$
c. lye, NaOH **d.** sugar, $C_{12}H_{22}O_{11}$

Chemistry Link to Industry

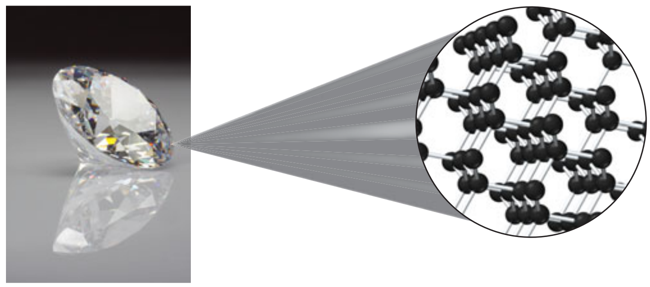
MANY FORMS OF CARBON

Carbon, which has the symbol C and atomic number 6, is located in Group 4A (14) in Period 2 on the periodic table. However, its atoms can be arranged in different ways to give several different substances. Two forms of carbon—diamond and graphite—have been known since prehistoric times. In diamond, carbon atoms are arranged in a rigid structure. A diamond is transparent and harder than any other substance, whereas graphite is black and soft. In graphite, carbon atoms are arranged in sheets that slide over each other. Graphite is used as pencil lead, as lubricants, and as carbon fibers for the manufacture of light weight golf clubs and tennis rackets.

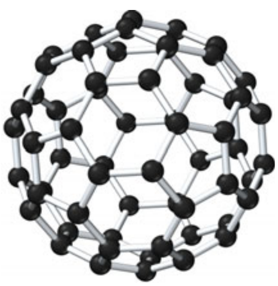
Two other forms of carbon have been discovered more recently. In the form called buckminsterfullerene or buckyball, 60 carbon atoms are arranged as rings of 5 and 6 atoms to give a spherical, cage-like structure. When a fullerene structure is stretched out, it produces a cylinder with a diameter of only a few nanometers called a nanotube. Practical uses for buckyballs and nanotubes are not yet developed, but they are expected to find use in light weight structural materials, heat conductors, computer parts, and medicine.



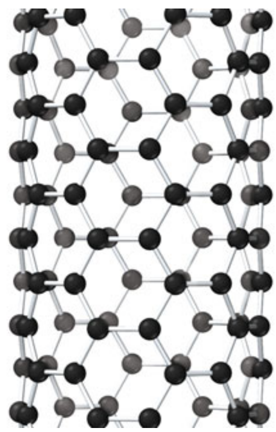
(a) Graphite



(b) Diamond



(c) Buckminsterfullerene



(d) Nanotubes

Carbon atoms can form different types of structures.

3.2 The Periodic Table

LEARNING GOAL

Use the periodic table to identify the group and the period of an element; identify the element as a metal, nonmetal, or metalloid.

As more and more elements were discovered, it became necessary to organize them into some type of classification system. By the late 1800s, scientists recognized that certain elements looked alike and behaved much the same way. In 1872, a Russian chemist, Dmitri Mendeleev, arranged the 60 elements known at that time into groups with similar properties and placed them in order of increasing atomic masses. Today, this arrangement of 117 elements is known as the **periodic table** (see Figure 3.1).

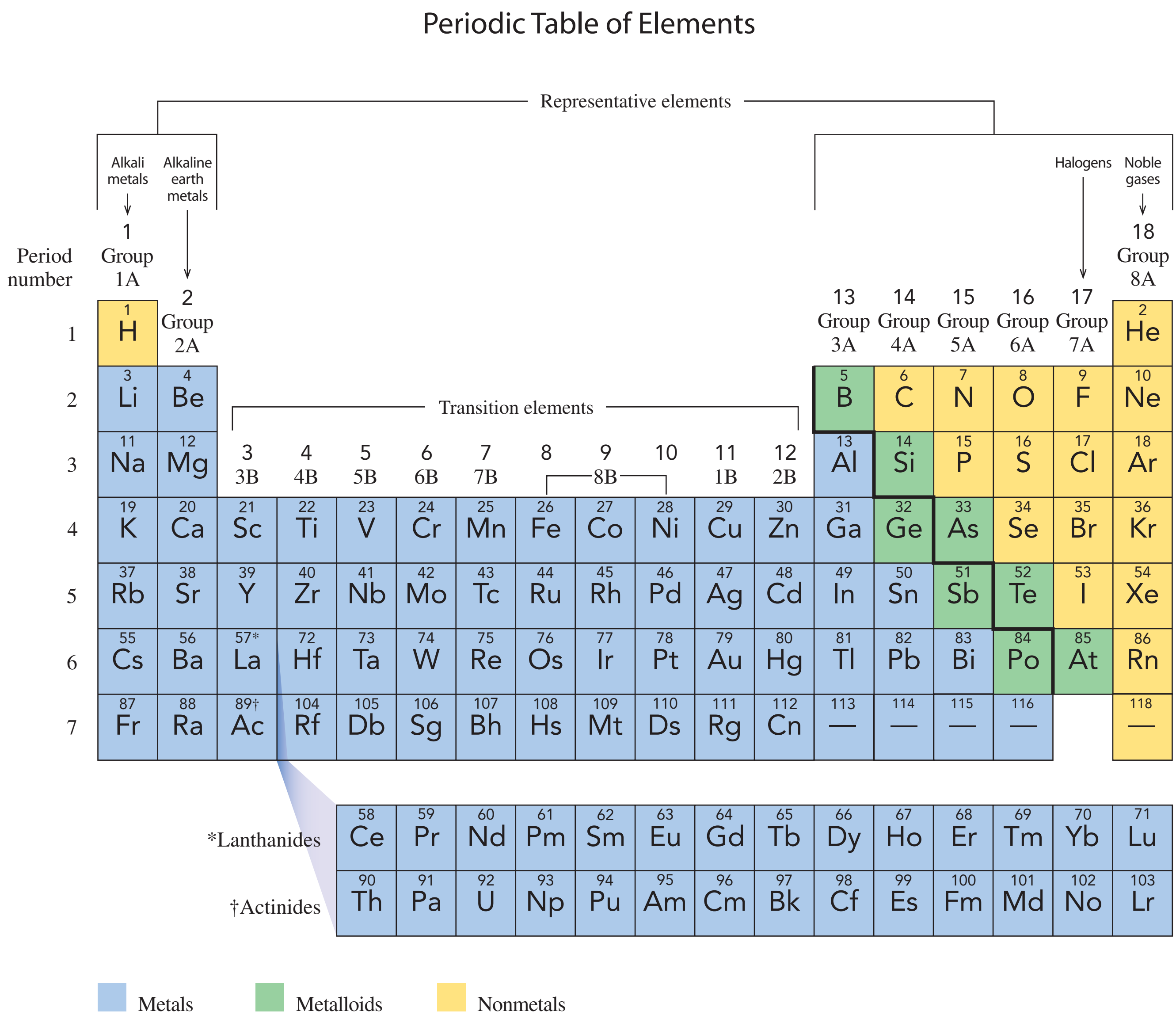


FIGURE 3.1 On the periodic table, groups are elements arranged as vertical columns, and periods are the elements in each horizontal row.

Q What is the symbol of the alkali metal in Period 3?

MC

TUTORIAL
Elements and Symbols in the Periodic Table

Periods and Groups

Each horizontal row in the periodic table is called a **period** (see Figure 3.2). Each period is counted from the top of the table as Period 1 to Period 7. The first period contains 2 elements: hydrogen (H) and helium (He). The second period contains 8 elements: lithium (Li), beryllium (Be), boron (B), carbon (C), nitrogen (N), oxygen (O), fluorine (F), and neon (Ne). The third period also contains 8 elements beginning with sodium (Na) and ending with argon (Ar). The fourth period, which begins with potassium (K), and the fifth period, which begins with rubidium (Rb), have 18 elements each. The sixth period, which begins with cesium (Cs), has 32 elements. The seventh period, as of today, contains the 31 remaining elements, although it could go up to 32.

Each vertical column on the periodic table contains a **group** (or family) of elements that have similar properties. At the top of each column is a number that is assigned to each group. The elements in the first two columns on the left and the last six columns on the right of the periodic table are called the **representative elements**. For many years, they have been given group numbers 1A–8A. In the center of the periodic table is a block

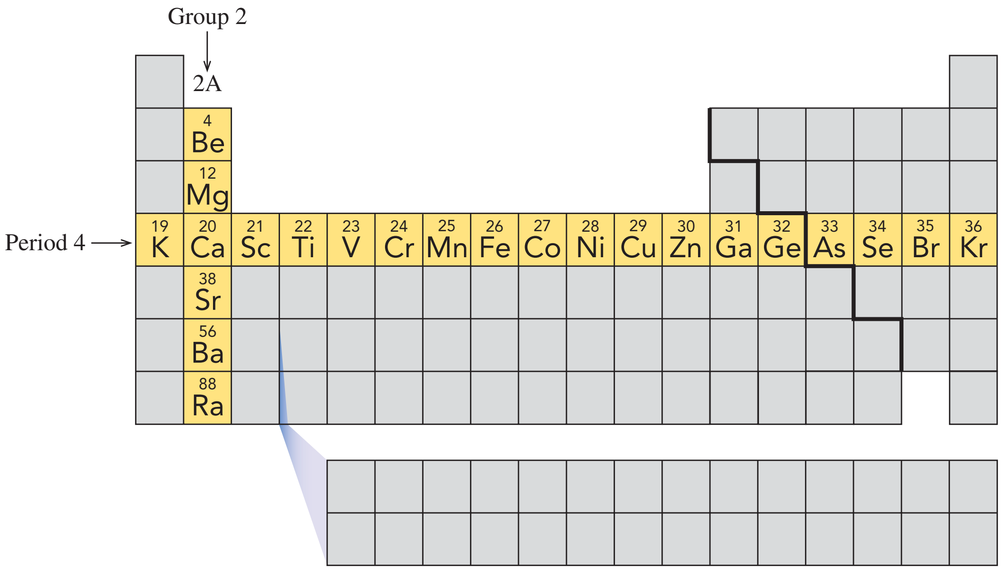


FIGURE 3.2 On the periodic table, each vertical column represents a group of elements and each horizontal row of elements represents a period.

Q Are the elements Si, P, and S part of a group or a period?

of elements known as the **transition elements**, which are designated with the letter "B." A newer numbering system assigns group numbers of 1–18 going across the periodic table. Because both systems of group numbers are currently in use, they are both indicated on the periodic table in this text and are included in our discussions of elements and group numbers. The lanthanides and actinides that are part of Periods 6 and 7 are placed at the bottom of the periodic table to allow it fit on a page.

Classification of Groups

Several groups in the periodic table have special names (see Figure 3.3). Group 1A (1) elements—lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr)—are a family of elements known as the **alkali metals** (see Figure 3.4). The elements within this group are soft, shiny metals that are good conductors of heat and electricity and have relatively low melting points. Alkali metals react vigorously with water and form white products when they combine with oxygen.

Although hydrogen (H) is at the top of Group 1A (1), hydrogen is not an alkali metal and has very different properties than the rest of the elements in this group. Thus hydrogen is not included in the classification of alkali metals.

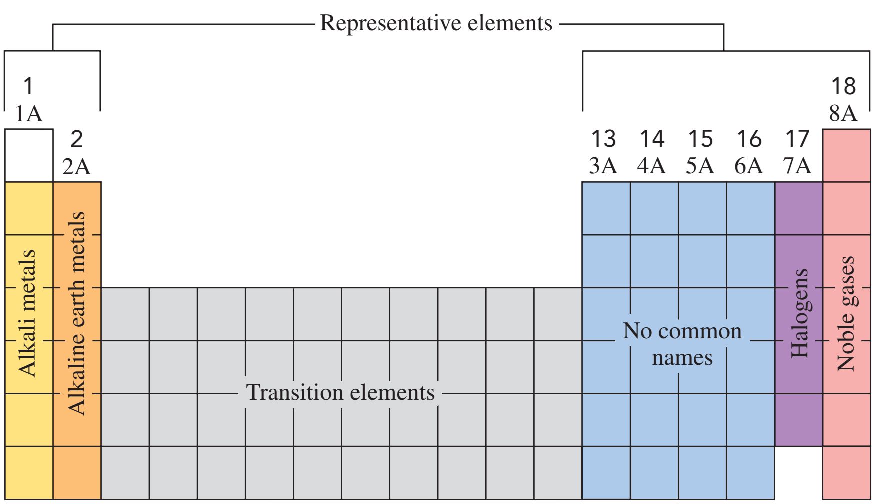


FIGURE 3.3 Certain groups on the periodic table have common names.

Q What is the common name for the group of elements that includes helium and argon?

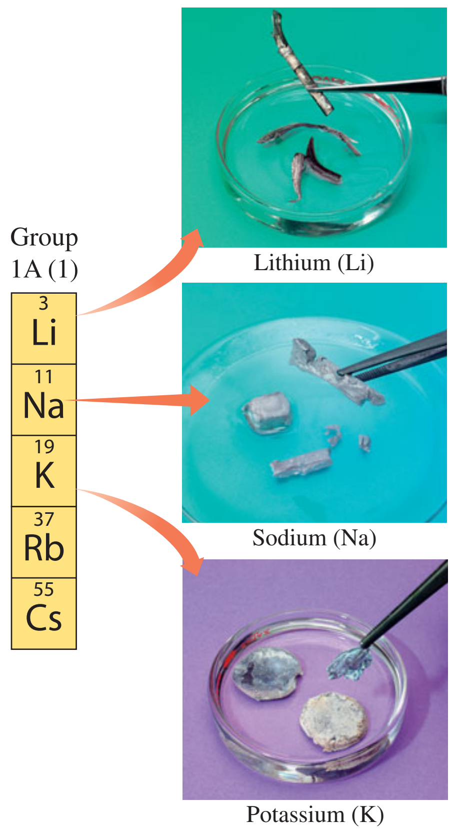


FIGURE 3.4 Lithium (Li), sodium (Na), and potassium (K) are some alkali metals from Group 1A (1).

Q What physical properties do these alkali metals have in common?

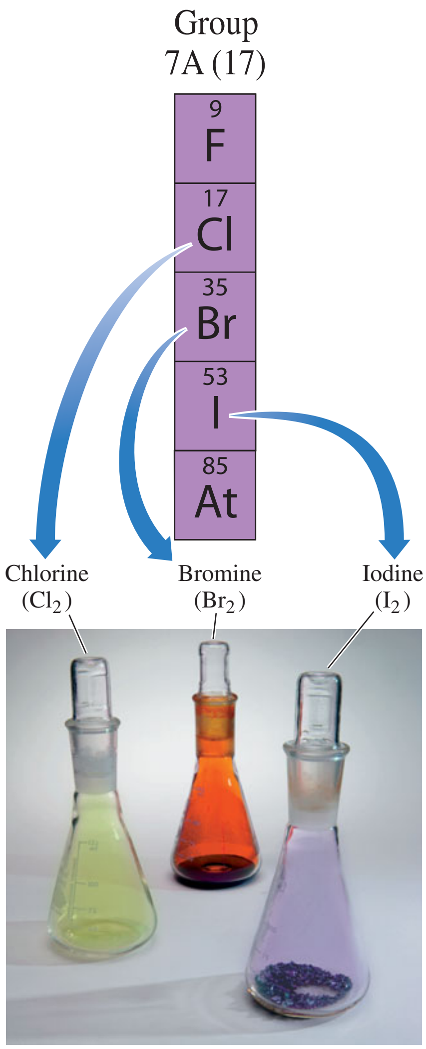


FIGURE 3.5 Chlorine (Cl_2), bromine (Br_2), and iodine (I_2) are examples of halogens from Group 7A (17).

Q What elements are in the halogen group?

Group 2A (2) elements—beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra)—are called the **alkaline earth metals**. They are shiny metals like those in Group 1A (1), but they are not as reactive.

The **halogens** are found on the right side of the periodic table in Group 7A (17). They include the elements fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At) (see Figure 3.5). The halogens, especially fluorine and chlorine, are highly reactive and form compounds with most of the elements.

Group 8A (18) contains the **noble gases**—helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). They are quite unreactive and are seldom found in combination with other elements.

SAMPLE PROBLEM 3.3

Group and Period Numbers of Some Elements

Give the period and group for each of the following elements and identify as a representative or transition element:

a. iodine **b.** manganese **c.** barium **d.** gold

SOLUTION

a. Iodine (I), Period 5, Group 7A (17), is a representative element.
b. Manganese (Mn), Period 4, Group 7B (7), is a transition element.
c. Barium (Ba), Period 6, Group 2A (2), is a representative element.
d. Gold (Au), Period 6, Group 1B (11), is a transition element.

STUDY CHECK 3.3

Strontium is an element that gives a brilliant red color to fireworks.

a. In what group is strontium found?
b. In what chemical family is strontium found?
c. In what period is strontium found?
d. What are the name and symbol of the element in Period 3 that is in the same group as strontium?
e. What alkali metal, halogen, and noble gas are in the same period as strontium?

Strontium provides the red color in fireworks.

Chemistry Link to Health

ELEMENTS ESSENTIAL TO HEALTH

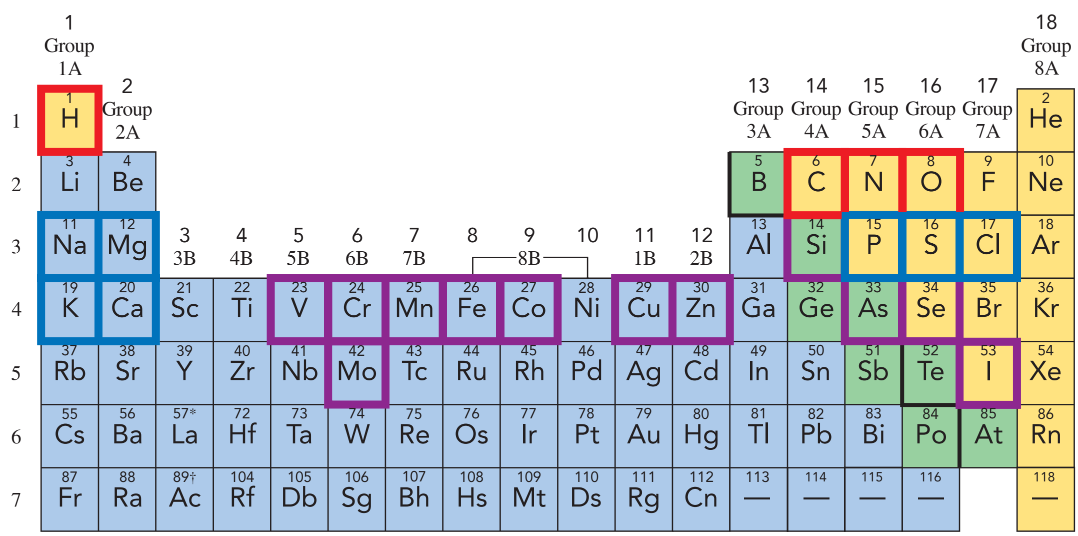
Of all the elements, only about 20 are essential for the well-being and survival of the human body. Of those, four elements—oxygen, carbon, hydrogen, and nitrogen—which are representative elements in Period 1 and Period 2 on the periodic table, make up 96% of our body mass. Most of the food in our daily diet provides these elements to maintain a healthy body. These elements are found in carbohydrates, fats, and proteins. Most of the hydrogen and oxygen is found in water, which makes up 55–60% of our body mass.

The macrominerals—Ca, P, K, Cl, S, Na, and Mg—are located in Period 3 and Period 4 of the periodic table. They are involved in the formation of bones and teeth, maintenance of heart and blood vessels, muscle contraction, nerve impulses, acid–base balance of body fluids, and regulation of cellular metabolism. The macrominerals are present in lower amounts than the major elements, so that smaller amounts are required in our daily diets.

The other essential elements, called microminerals or trace elements, are mostly transition elements in Period 4 along with Mo and I in Period 5. They are present in the human body in small amounts, some less than 100 mg. In recent years, the detection of such small amounts has improved so that researchers can more easily identify the roles of trace elements. Some trace elements such as arsenic, chromium and selenium are toxic at higher levels in the body but are still required by the body. Other elements such as tin and nickel are thought to be essential, but their metabolic role has not yet been determined. Some examples and the amounts present in a 60-kg person are listed in Table 3.3.

TABLE 3.3 Typical Amounts of Essential Elements in a 60-kg Adult

Element	Quantity	Function
Major Elements		
Oxygen (O)	39 kg	Building block of biomolecules and water (H_2O)
Carbon (C)	11 kg	Building block of organic and biomolecules
Hydrogen (H)	6 kg	Component of biomolecules, water (H_2O), and pH of body fluids, stomach acid (HCl)
Nitrogen (N)	1.5 kg	Component of proteins and nucleic acids
Macrominerals		
Calcium (Ca)	1000 g	Bone and teeth, muscle contraction, nerve impulses
Phosphorus (P)	600 g	Bone and teeth, nucleic acids, ATP
Potassium (K)	120 g	Most abundant positive ion (K^+) in cells, muscle contraction, nerve impulses
Chlorine (Cl)	100 g	Most abundant negative ion (Cl^-) in fluids outside cells, stomach acid (HCl)
Sulfur (S)	86 g	Proteins, liver, vitamin B_1, insulin
Sodium (Na)	60 g	Most abundant positive ion (Na^+) in fluids outside cells, water balance, muscle contraction, nerve impulses
Magnesium (Mg)	36 g	Bone, required for metabolic reactions
Microminerals (trace elements)		
Iron (Fe)	3600 mg	Component of oxygen carrier hemoglobin
Silicon (Si)	3000 mg	Growth and maintenance of bone and teeth, tendons and ligaments, hair and skin
Zinc (Zn)	2000 mg	Metabolic reactions in cells, DNA synthesis, growth of bone, teeth, connective tissue, immune system
Copper (Cu)	240 mg	Blood vessels, blood pressure, immune system
Manganese (Mn)	60 mg	Bone growth, blood clotting, necessary for metabolic reactions
Iodine (I)	20 mg	Proper thyroid function
Molybdenum (Mo)	12 mg	Needed to process Fe and N from diets
Arsenic (As)	3 mg	Growth and reproduction
Chromium (Cr)	3 mg	Maintenance of blood sugar levels, synthesis of biomolecules
Cobalt (Co)	3 mg	Vitamin B_{12}, red blood cells
Selenium (Se)	2 mg	Immune system, health of heart and pancreas
Vanadium (V)	2 mg	Formation of bone and teeth, energy from food



Major elements in human body   Macrominerals   Microminerals (trace elements)

Metals, Nonmetals, and Metalloids

Another feature of the periodic table is the heavy zigzag line that separates the elements into the *metals* and the *nonmetals*. The metals are those elements on the left of the line *except for hydrogen*, and the nonmetals are the elements on the right (see Figure 3.6).

TUTORIAL
Metals, Nonmetals, and Metalloids

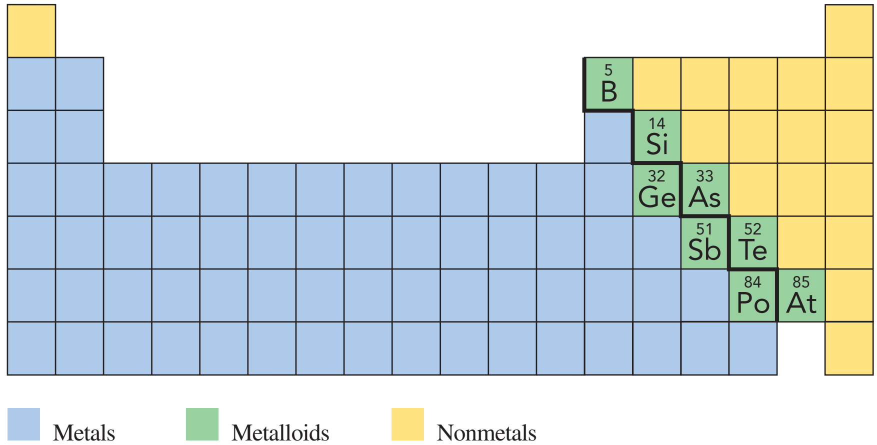


FIGURE 3.6 Along the heavy zigzag line on the periodic table that separates the metals and nonmetals are metalloids, which exhibit characteristics of both metals and nonmetals.
Q On which side of the heavy zigzag line are the nonmetals located?

In general, most **metals** are shiny solids, such as copper (Cu), gold (Au), and silver (Ag). Metals can be shaped into wires (ductile) or hammered into a flat sheet (malleable). Metals are good conductors of heat and electricity. They usually melt at higher temperatures than nonmetals. All the metals are solids at room temperature, except for mercury (Hg), which is a liquid.

Nonmetals are not especially shiny, ductile, or malleable, and they are often poor conductors of heat and electricity. They typically have low melting points and low densities. Some examples of nonmetals are hydrogen (H), carbon (C), nitrogen (N), oxygen (O), chlorine (Cl), and sulfur (S).

Except for aluminum, the elements located along the heavy line are **metalloids**: B, Si, Ge, As, Sb, Te, Po, and At. Metalloids are elements that exhibit some properties that are typical of the metals and other properties that are characteristic of the nonmetals. For example, they are better conductors of heat and electricity than the nonmetals, but not as good as the metals. The metalloids are semiconductors because they can be modified to function as conductors or insulators. Table 3.4 compares some characteristics of silver, a metal, with those of antimony, a metalloid, and sulfur, a nonmetal.

A silver cup is shiny, antimony is a blue-gray solid, and sulfur is a dull, yellow color.

TABLE 3.4 Some Characteristics of a Metal, a Metalloid, and a Nonmetal

Silver (Ag)	Antimony (Sb)	Sulfur (S)
Metal	Metalloid	Nonmetal
Shiny	Blue-gray, shiny	Dull, yellow
Extremely ductile	Brittle	Brittle
Can be hammered into sheets (malleable)	Shatters when hammered	Shatters when hammered
Good conductor of heat and electricity	Poor conductor of heat and electricity	Poor conductor of heat and electricity, good insulator
Used in coins, jewelry, tableware	Used to harden lead, color glass and plastics	Used in gunpowder, rubber, fungicides
Density 10.5 g/mL	Density 6.7 g/mL	Density 2.1 g/mL
Melting point 962 °C	Melting point 630 °C	Melting point 113 °C

CONCEPT CHECK 3.2

Groups and Periods on the Periodic Table

Consider the elements aluminum, germanium, and phosphorus.

a. In what group and period are they found?
b. Identify each as a metal, nonmetal, or metalloid.

ANSWER

a. Aluminum is in Group 3A (13), Period 3. Germanium is in Group 4A (14), Period 4. Phosphorus is in Group 5A (15), Period 3.
b. Aluminum is a metal, germanium is a metalloid, and phosphorus is a nonmetal.

SAMPLE PROBLEM 3.4

Metals, Nonmetals, or Metalloids

Use the periodic table to classify each of the following elements by its group and period, group name (if any), and as a metal, nonmetal, or metalloid:

a. Na **b.** I **c.** B

SOLUTION

a. Na (sodium), Group 1A (1), Period 3, is an alkali metal.
b. I (iodine), Group 7A (17), Period 5, halogen, is a nonmetal.
c. B (boron), Group 3A (13), Period 2, is a metalloid.

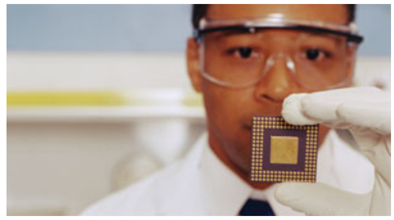
STUDY CHECK 3.4

Give the name and symbol of the following elements:

a. Group 5A (15), Period 4
b. A noble gas in Period 3
c. A metalloid in Period 3

Career Focus

MATERIALS SCIENTIST



"The unique qualities of semiconducting metals make it possible for us to create sophisticated electronic circuits," says Tysen Streib, Global Product Manager, Applied Materials. "Elements from Groups 3A (13), 4A (14), and 5A (15) of the periodic table make good semiconductors because they readily form covalently bonded crystals. When small amounts of impurities are added, free-flowing electrons or holes can travel through the crystal with very little interference. Without these covalent bonds and loosely bound electrons, we wouldn't have any of the microchips that we use in computers, cell phones, and thousands of other devices."

Materials scientists study the chemical properties of materials to find new uses for them in products such as cars, bridges, and clothing. They also develop materials that can be used as superconductors or in integrated-circuit chips and fuel cells. Chemistry is important in materials science because it provides information about structure and composition.

QUESTIONS AND PROBLEMS

The Periodic Table

3.7 Identify the group or period number described by each of the following:
a. contains the elements C, N, and O
b. begins with helium
c. the alkali metals
d. ends with neon

3.8 Identify the group or period number described by each of the following:
a. contains Na, K, and Rb
b. the row that begins with Li
c. the noble gases
d. contains F, Cl, Br, and I

3.9 Classify each of the following as an alkali metal, alkaline earth metal, transition element, halogen, or noble gas:
a. Ca **b.** Fe **c.** Xe
d. K **e.** Cl

3.10 Classify each of the following as an alkali metal, alkaline earth metal, transition element, halogen, or noble gas:
a. Ne **b.** Mg **c.** Cu
d. Br **e.** Ba

3.11 Give the symbol of the element described by each of the following:
a. Group 4A (14), Period 2
b. a noble gas in Period 1
c. an alkali metal in Period 3
d. Group 2A (2), Period 4
e. Group 3A (13), Period 3

3.12 Give the symbol of the element described by each of the following:
a. an alkaline earth metal in Period 2
b. Group 5A (15), Period 3
c. a noble gas in Period 4
d. a halogen in Period 5
e. Group 4A (14), Period 4

3.13 Is each of the following elements a metal, nonmetal, or metalloid?
a. calcium
b. sulfur
c. a shiny element
d. an element that is a gas at room temperature
e. located in Group 8A (18)
f. bromine
g. tellurium
h. silver

3.14 Is each of the following elements a metal, nonmetal, or metalloid?
a. located in Group 2A (2)
b. a good conductor of electricity
c. chlorine
d. arsenic
e. an element that is not shiny
f. oxygen
g. nitrogen
h. tin

LEARNING GOAL

Describe the electrical charge and location in an atom for a proton, a neutron, and an electron.

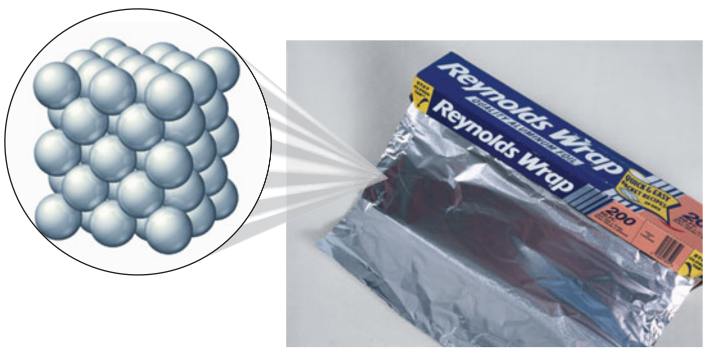
SELF STUDY ACTIVITY
Atoms and Isotopes

TUTORIAL
The Anatomy of Atoms

3.3 The Atom

All the elements listed on the periodic table are made up of atoms. In Chapter 2, we described an **atom** as the smallest particle of an element that retains the characteristics of that element. Imagine that you are tearing a piece of aluminum foil into smaller and smaller pieces. Now imagine that you have a piece so small that you cannot tear it apart any further. Then you would have a single atom of aluminum.

The concept of the atom is relatively recent. Although the Greek philosophers in 500 B.C.E. reasoned that everything must contain minute particles they called *atomos*, the idea of atoms did not become a scientific theory until 1808. Then John Dalton (1766–1844) developed an atomic theory that proposed that atoms were responsible for the combinations of elements found in compounds.



Aluminum foil consists of atoms of aluminum.

Dalton's Atomic Theory

1. All matter is made up of tiny particles called atoms.
2. All atoms of a given element are similar to one another and different from atoms of other elements.
3. Atoms of two or more different elements combine to form compounds. A particular compound is always made up of the same kinds of atoms and always has the same number of each kind of atom.
4. A chemical reaction involves the rearrangement, separation, or combination of atoms. Atoms are never created or destroyed during a chemical reaction.

Dalton's atomic theory formed the basis of current atomic theory, although we have modified some of Dalton's statements. We now know that atoms of the same element are not completely identical to each other and consist of even smaller particles. However, an atom is still the smallest particle that retains the properties of an element.

Although atoms are the building blocks of everything we see around us, we cannot see an atom or even a billion atoms with the naked eye. However, when billions and billions of atoms are packed together, the characteristics of each atom are added to those of the next until we can see the characteristics we associate with the element. For example, a small piece of the shiny element nickel consists of many, many nickel atoms. A special kind of microscope called a *scanning tunneling microscope* (STM) produces images of individual atoms (see Figure 3.7).

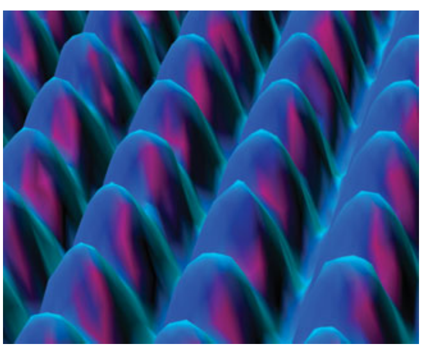


FIGURE 3.7 Images of nickel atoms are produced when nickel is magnified millions of times by a scanning tunneling microscope (STM).
Q Why is a microscope with extremely high magnification needed to see these atoms?

Electrical Charges in an Atom

By the end of the 1800s, experiments with electricity showed that atoms were not solid spheres but were composed of even smaller bits of matter called *subatomic particles*, three of which are the *proton*, *neutron*, and *electron*. Two of these subatomic particles were discovered because they have electrical charges.

An electrical charge can be positive or negative. Experiments show that like charges repel, or push away from each other. When you brush your hair on a dry day, electrical charges that are alike build up on the brush and in your hair. As a result, your hair flies away from the brush. However, opposite or unlike charges attract. The clinginess of the clothing taken from the clothes dryer is due to the attraction of opposite, unlike charges, as shown in Figure 3.8.

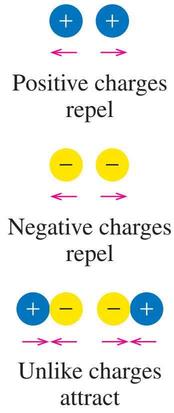


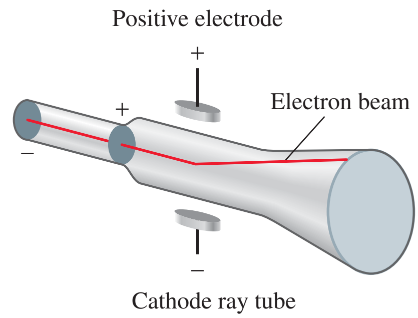
FIGURE 3.8 Like charges repel and unlike charges attract.
Q Why are the electrons attracted to the protons in the nucleus of an atom?

Structure of the Atom

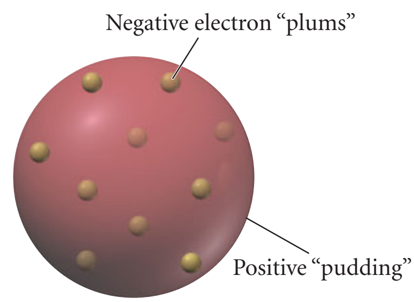
In 1897, J. J. Thomson, an English physicist, applied electricity to a glass tube, which produced streams of small particles called *cathode rays*. Because these rays were attracted to a positively charged electrode, Thomson realized that the particles in the rays must be negatively charged. In further experiments, these particles called **electrons** were found to be much smaller than the atom and to have extremely small masses. Because atoms are neutral, scientists soon discovered that atoms contained positively charged particles called **protons** that were much heavier than the electrons.

Thomson proposed the “plum-pudding” model for the atom in which the electrons and protons were randomly distributed through the atom. In 1911, Ernest Rutherford worked with Thomson to test this model. In Rutherford’s experiment, positively charged particles were aimed at a thin sheet of gold foil (see Figure 3.9). If the Thomson model were correct, the particles would travel in straight paths through the gold foil. Rutherford was greatly surprised to find that some of the particles were deflected as they passed through the gold foil, and a few particles were deflected so much that they went back in the opposite direction. According to Rutherford, it was as though he had shot a cannonball at a piece of tissue paper, and it bounced back at him.

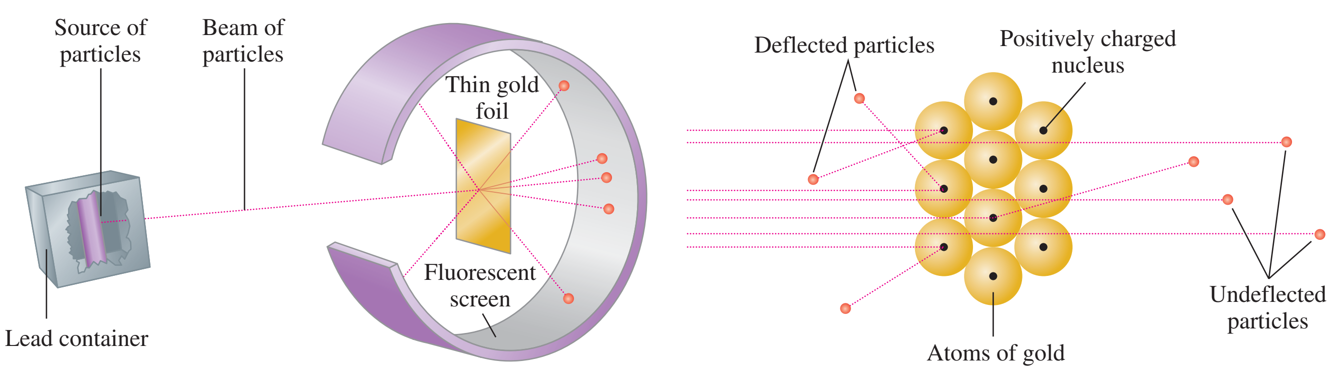
From the gold-foil experiments, Rutherford realized that the protons must be contained in a small, positively charged region at the center of the atom, which he called the **nucleus**. He proposed that the electrons in the atom occupy the space surrounding the nucleus through which most of the particles traveled undisturbed. Only the particles



Negatively charged cathode rays (electrons) are attracted to the positive electrode.



Thomson proposed the “plum-pudding” model of the atom.



(a) **(b)**

FIGURE 3.9 **(a)** Positive particles are aimed at a piece of gold foil. **(b)** Particles that come close to the atomic nuclei are deflected from their straight path.
Q Why are some particles deflected while most pass through the gold foil undeflected?

that came near this dense, positive center were deflected. If an atom were the size of a football stadium, the nucleus would be about the size of a golf ball placed in the center of the field.

Scientists knew that the nucleus was heavier than the mass of the protons, so they looked for another subatomic particle. Eventually, they discovered that the nucleus also contained a particle that is neutral, which they called a **neutron**. Thus, the masses of the protons and neutrons in the nucleus determine the mass of an atom (see Figure 3.10).

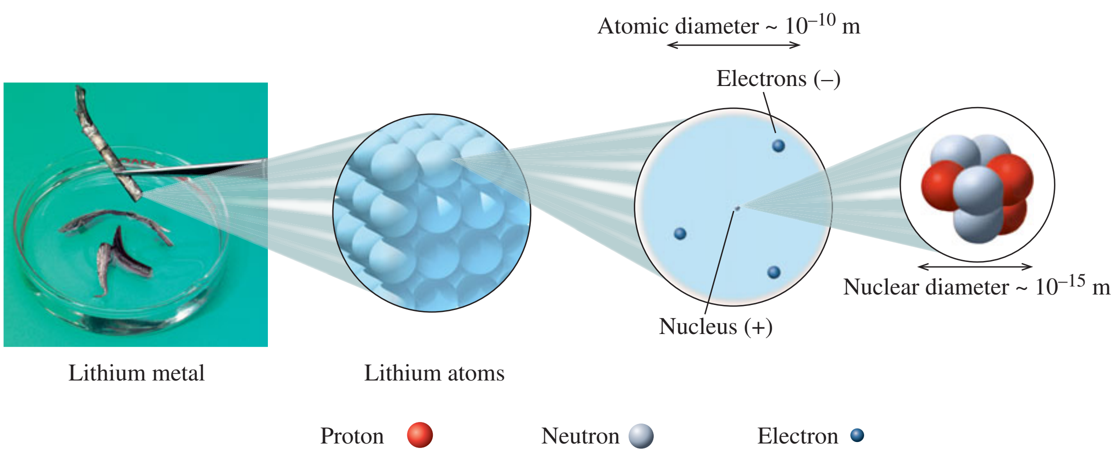


FIGURE 3.10 In an atom, the protons and neutrons that make up almost all the mass are packed into the tiny volume of the nucleus. The electrons surround the nucleus and account for the large volume of the atom.

Q Why can we say that the atom is mostly empty space?

TUTORIAL
Atomic Structure and Properties of Subatomic Particles

Mass of the Atom

All the subatomic particles are extremely small compared with the things you see around you. One proton has a mass of 1.7×10^{-24} g, and the neutron is about the same. However, the electron has a mass 9.1×10^{-28} g, which is about 1/2000th of the mass of either a proton or neutron. Because the masses of subatomic particles are so small, chemists use a very small unit of mass called an **atomic mass unit (amu)**. An amu is defined as one-twelfth of the mass of a carbon atom which has a nucleus containing six protons and six neutrons. In biology, the atomic mass unit is called a *Dalton* (Da) in honor of John Dalton. On the amu scale, the proton and neutron each have a mass of about 1 amu. Table 3.5 summarizes some information about the subatomic particles in an atom.

TABLE 3.5 Subatomic Particles in the Atom

Particle	Symbol	Charge	Mass (amu)	Location in Atom
Proton	p or p^+	1+	1.007	Nucleus
Neutron	n or n^0	0	1.008	Nucleus
Electron	e^-	1−	0.000 55	Outside nucleus

CONCEPT CHECK 3.3

Identifying Subatomic Particles

Is each of the following statements true or false?

a. Protons are heavier than electrons.
b. Protons are attracted to neutrons.
c. Electrons are so small that they have no electrical charge.
d. The nucleus contains all the protons and neutrons of an atom.

ANSWER

a. True
b. False; protons are attracted to electrons.
c. False; electrons have a 1− charge.
d. True

SAMPLE PROBLEM 3.5

Identifying Subatomic Particles

Identify the subatomic particle that has the following characteristics:

a. no charge
b. a mass of 0.000 55 amu
c. a mass about the same as a neutron

SOLUTION

a. neutron **b.** electron **c.** proton

STUDY CHECK 3.5

Is the following statement true or false?

The nucleus occupies a large volume in an atom.

Explore Your World

REPULSION AND ATTRACTION

Tear a small piece of paper into bits. Brush your hair several times, and place the brush just above the bits of paper. Use your knowledge of electrical charges to give an explanation for your observations. Try the same experiment with a comb.

QUESTIONS

1. What happens when objects with like charges are placed close together?
2. What happens when objects with unlike charges are placed close together?

QUESTIONS AND PROBLEMS

The Atom

3.15 Identify each of the following as describing either a proton, neutron, or electron:
a. has the smallest mass
b. has a 1+ charge
c. is found outside the nucleus
d. is electrically neutral

3.16 Identify each of the following as describing either a proton, neutron, or electron:
a. has a mass about the same as a proton
b. is found in the nucleus
c. is attracted to the protons
d. has a 1− charge

3.17 What did Rutherford determine about the structure of the atom from his gold-foil experiment?

3.18 Why does the nucleus in every atom have a positive charge?

3.19 Identify each of the following statements as true or false:
a. A proton and an electron have opposite charges.
b. The nucleus contains most of the mass of an atom.
c. Electrons repel each other.
d. A proton is attracted to a neutron.

3.20 Identify each of the following statements as true or false:
a. A proton is attracted to an electron.
b. A neutron has twice the mass of a proton.
c. Neutrons repel each other.
d. Electrons and neutrons have opposite charges.

3.21 On a dry day, your hair flies away when you brush it. How would you explain this?

3.22 Sometimes clothes removed from the dryer cling together. What kinds of charges are on the clothes?

3.4 Atomic Number and Mass Number

LEARNING GOAL

Given the atomic number and the mass number of an atom, state the number of protons, neutrons, and electrons.

All the atoms of the same element always have the same number of protons. This feature distinguishes atoms of one element from atoms of all the other elements.

Atomic Number

An **atomic number**, which is equal to the number of protons in the nucleus of an atom, is used to identify and define each element.

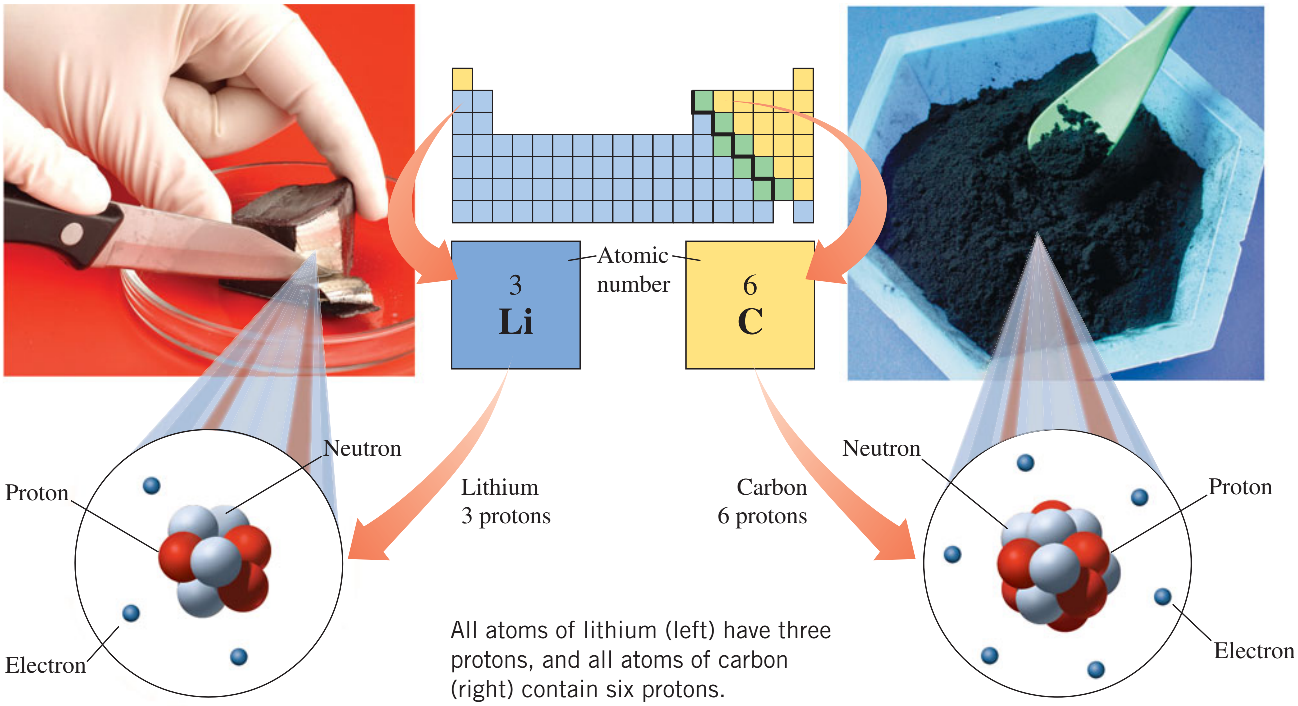
Atomic number = number of protons in an atom

On the inside front cover of this text is a periodic table, which gives all the elements in order of increasing atomic number. The atomic number is the whole number that appears above the symbol of each element. For example, a hydrogen atom, with atomic

TUTORIAL
Element Names, Symbols, and Atomic Numbers

number 1, has 1 proton; a lithium atom, with atomic number 3, has 3 protons; an atom of carbon, with atomic number 6, has 6 protons; and gold, with atomic number 79, has 79 protons; and so forth.

An atom is electrically neutral. That means that the number of protons in an atom is equal to the number of electrons. This electrical balance gives an atom an overall charge of zero. Thus, in every atom, the atomic number also gives the number of electrons.



All atoms of lithium (left) have three protons, and all atoms of carbon (right) contain six protons.

SAMPLE PROBLEM 3.6

Using Atomic Number to Find the Number of Protons and Electrons

Using the periodic table in Figure 3.1, state the atomic number, number of protons, and number of electrons for an atom of each of the following elements:

a. nitrogen **b.** magnesium **c.** bromine

SOLUTION

a. atomic number 7; 7 protons and 7 electrons
b. atomic number 12; 12 protons and 12 electrons
c. atomic number 35; 35 protons and 35 electrons

STUDY CHECK 3.6

Consider an atom that has 79 electrons.

a. How many protons are in its nucleus?
b. What is its atomic number?
c. What is its name, and what is its symbol?

Mass Number

MC TUTORIAL
Atomic Number and Mass Number

We now know that the protons and neutrons determine the mass of the nucleus. For any atom, the **mass number** is the sum of the number of protons and neutrons in the nucleus of a single atom. Thus, the mass number is a counting number, which is always a whole number. Because mass number represents the particles in the nucleus of a single atom, it does not appear on the periodic table.

Mass number = number of protons + number of neutrons

For example, an atom of oxygen that contains 8 protons and 8 neutrons has a mass number of 16. An atom of iron that contains 26 protons and 30 neutrons has a mass number of 56. Table 3.6 illustrates the relationship between atomic number, mass number, and the number of protons, neutrons, and electrons in examples of atoms for different elements.

TABLE 3.6 Composition of Some Atoms of Different Elements

Element	Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
Hydrogen	H	1	1	1	0	1
Nitrogen	N	7	14	7	7	7
Chlorine	Cl	17	37	17	20	17
Iron	Fe	26	57	26	31	26
Gold	Au	79	197	79	118	79

CONCEPT CHECK 3.4

Counting Subatomic Particles in Atoms

An atom of silver has a mass number of 109.

a. How many protons are in the nucleus?
b. How many neutrons are in the nucleus?
c. How many electrons are in the atom?

ANSWER

a. Silver (Ag) with atomic number 47 has 47 protons.
b. The number of neutrons is calculated by subtracting the number of protons from the mass number. 109 − 47 = 62 neutrons for Ag with a mass number of 109.
c. An atom is neutral, which means that the number of electrons is equal to the number of protons. An atom of silver with 47 protons has 47 electrons.

SAMPLE PROBLEM 3.7

Calculating Numbers of Protons, Neutrons, and Electrons

For an atom of zinc that has a mass number of 68, determine the following:

a. the number of protons
b. the number of neutrons
c. the number of electrons

SOLUTION

a. Zinc (Zn), with an atomic number of 30, has 30 protons.
b. The number of neutrons in this atom is found by subtracting the atomic number from the mass number.

$$\begin{array}{ccccc} \text{Mass number} & - & \text{atomic number} & = & \text{number of neutrons} \\ 68 & - & 30 & = & 38 \end{array}$$

c. Because the zinc atom is neutral, the number of electrons is equal to the number of protons. A zinc atom has 30 electrons.

STUDY CHECK 3.7

How many neutrons are in the nucleus of a bromine atom that has a mass number of 80?

Career Focus

OPTICIAN

"When a patient brings in a prescription, I help select the proper lenses, put them into a frame, and fit them properly on the patient's face," says Suranda Lara, optician, Kaiser Hospital. "If a prescription requires a thinner and lighter-weight lens, we formulate that lens. So we have to understand the different materials used to make lenses. Sometimes patients come in with their own glasses that they want to convert to sunglasses. We remove the lenses and put them into a tint bath, which turns them into sunglasses." Opticians fit and adjust eyewear for patients who have had their eyesight tested by an ophthalmologist or optometrist. Optics and mathematics are used to select materials for frames and lenses that are compatible with patients' facial measurements and lifestyles.

QUESTIONS AND PROBLEMS

Atomic Number and Mass Number

3.23 Would you use the atomic number, mass number, or both to determine each of the following?
a. number of protons in an atom
b. number of neutrons in an atom
c. number of particles in the nucleus
d. number of electrons in a neutral atom

3.24 What do you know about the subatomic particles from each of the following?
a. atomic number
b. mass number
c. mass number − atomic number
d. mass number + atomic number

3.25 Write the names and symbols of the elements with the following atomic numbers:
a. 3 **b.** 9 **c.** 20 **d.** 30
e. 10 **f.** 14 **g.** 53 **h.** 8

3.26 Write the names and symbols of the elements with the following atomic numbers:
a. 1 **b.** 11 **c.** 19 **d.** 82
e. 35 **f.** 47 **g.** 15 **h.** 2

3.27 How many protons and electrons are there in a neutral atom of each of the following elements?
a. argon
b. zinc
c. iodine
d. potassium

3.28 How many protons and electrons are there in a neutral atom of each of the following elements?
a. carbon **b.** fluorine
c. calcium **d.** sulfur

3.29 Complete the following table for a neutral atom of each element:

Name of the Element	Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
	Al		27			
		12			12	
Potassium					20	
				16	15	
			56			26

3.30 Complete the following table for a neutral atom of each element:

Name of the Element	Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
	N		15			
Calcium			42			
				38	50	
		14			16	
		56	138			

LEARNING GOAL

Give the number of protons, electrons, and neutrons in an isotope of an element; calculate the atomic mass of an element using the abundance and mass of its naturally occurring isotopes.

SELF STUDY ACTIVITY
Atoms and Isotopes

TUTORIAL
Isotopes

3.5 Isotopes and Atomic Mass

We have seen that all atoms of the same element have the same number of protons and electrons. However, the atoms of any one element are not completely identical because they can have different numbers of neutrons.

Isotopes are atoms of the same element that have the same number of protons, but different numbers of neutrons. For example, in a large sample of magnesium atoms, there are three different types of atoms or isotopes. We already know that all the isotopes of magnesium have 12 protons. However, one isotope has 12 neutrons, another has 13 neutrons, and yet another isotope has 14 neutrons. Because the mass number of a single atom is the sum of the protons and neutrons, the three isotopes of magnesium, which have the same atomic number, will have different mass numbers.

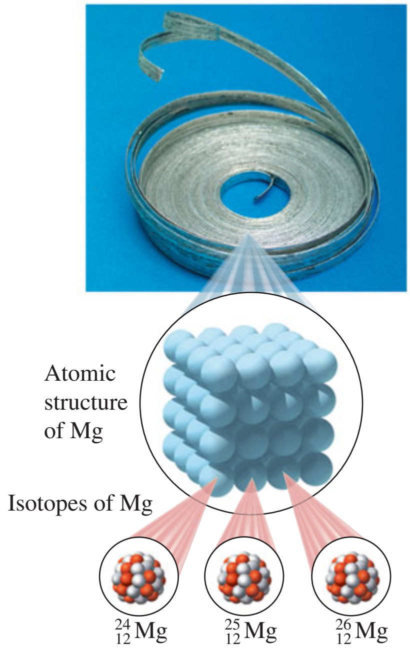
Atomic Symbols for the Isotopes of Magnesium

To distinguish between the different isotopes of the same element, we can write an **atomic symbol** that indicates the mass number of each isotope in the upper left corner and the atomic number of the element in the lower left corner.

Mass number $\rightarrow$ $\boxed{^{24}_{12}\text{Mg}}$ $\leftarrow$ Symbol of element
Atomic number $\rightarrow$

Atomic symbol for an isotope of magnesium.

An isotope may be referred to by its name or symbol followed by the mass number, such as magnesium-24 or Mg-24. Magnesium has three naturally occurring isotopes, as shown in Table 3.7. In a large sample of naturally occurring magnesium atoms, each type of isotope can be present as a low percentage or a high percentage. For example, the Mg-24 isotope makes up almost 80% of the total sample, whereas Mg-25 and Mg-26 each make up only about 10% of the total number of magnesium atoms.



The nuclei of three naturally occurring magnesium isotopes have different numbers of neutrons.

TABLE 3.7 Isotopes of Magnesium

Atomic Symbol	$^{24}_{12}$Mg	$^{25}_{12}$Mg	$^{26}_{12}$Mg
Number of protons	12	12	12
Number of electrons	12	12	12
Mass number	**24**	**25**	**26**
Number of neutrons	**12**	**13**	**14**
Mass of isotope (amu)	23.99	24.99	25.98
% abundance	78.70%	10.13%	11.17%

SAMPLE PROBLEM 3.8

Identifying Protons and Neutrons in Isotopes

State the number of protons and neutrons for each of the following isotopes of neon (Ne):

a. $^{20}_{10}$Ne **b.** $^{21}_{10}$Ne **c.** $^{22}_{10}$Ne

SOLUTION

The atomic number of Ne is 10, which means that the nucleus of each isotope has 10 protons. The number of neutrons in each isotope is found by subtracting the atomic number (10) from each of their mass numbers.

a. 10 protons; 10 neutrons (20 – 10)
b. 10 protons; 11 neutrons (21 – 10)
c. 10 protons; 12 neutrons (22 – 10)

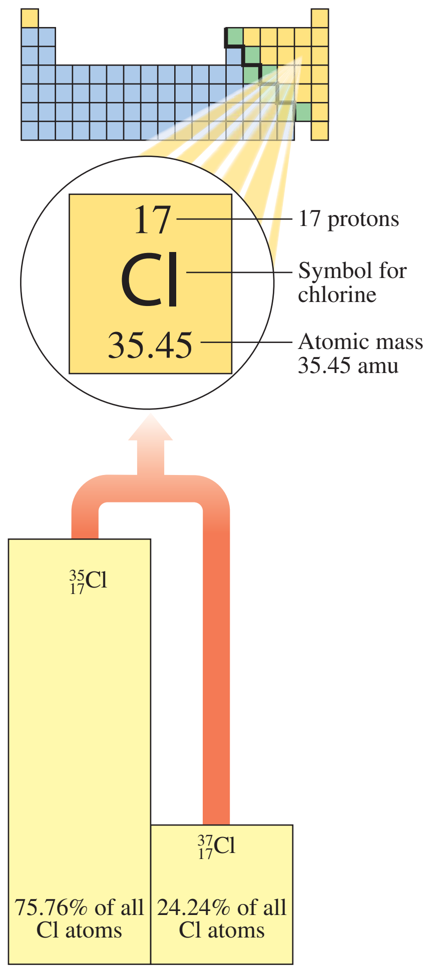
STUDY CHECK 3.8

Write an atomic symbol for each of the following:

a. a nitrogen atom with 8 neutrons
b. an atom with 20 protons and 22 neutrons
c. an atom with mass number 27 and 14 neutrons

Atomic Mass

In laboratory work, a chemist generally uses samples with many atoms that contain all the different atoms or isotopes of an element. Because each isotope has a different mass, chemists have calculated an **atomic mass** for an "average atom," which is a *weighted average* of the masses of all the naturally occurring isotopes of that element. On the periodic table, the atomic mass is the number including decimal places that is given below the symbol of each element. Most elements consist of two or more isotopes, which is one reason that the atomic masses on the periodic table are seldom whole numbers.



Chlorine, with two naturally occurring isotopes, has an atomic mass of 35.45 amu.

TUTORIAL
Atomic Mass Calculations

Calculating Atomic Mass

To calculate the atomic mass of an element, the percentage abundance of each isotope and its mass must be determined experimentally. For example, a large sample of naturally occurring chlorine atoms consists of 75.76% of ${}^{35}_{17}Cl$ atoms and 24.24% of ${}^{37}_{17}Cl$ atoms. The atomic mass is calculated using the percentage of each isotope and its mass: The ${}^{35}_{17}Cl$ isotope has a mass of 34.97 amu and the ${}^{37}_{17}Cl$ isotope has a mass of 36.97 amu.

$$\text{Atomic mass of Cl} = \underbrace{\text{mass of } {}^{35}_{17}\text{Cl} \times \frac{{}^{35}_{17}\text{Cl}\%}{100\%}}_{\text{amu from } {}^{35}_{17}\text{Cl}} + \underbrace{\text{mass of } {}^{37}_{17}\text{Cl} \times \frac{{}^{37}_{17}\text{Cl}\%}{100\%}}_{\text{amu from } {}^{37}_{17}\text{Cl}}$$

Isotope	Mass (amu)	×	Abundance (%)	=	Contribution to Average Cl Atom
${}^{35}_{17}Cl$	34.97	×	$\frac{75.76}{100}$	=	26.49 amu
${}^{37}_{17}Cl$	36.97	×	$\frac{24.24}{100}$	=	8.962 amu
			Atomic mass of Cl	=	35.45 amu

The atomic mass of 35.45 amu is the weighted average mass of a sample of Cl atoms, although no individual Cl atom actually has this mass. An atomic mass of 35.45 for chlorine also indicates that there is a higher percentage of ${}^{35}_{17}Cl$ atoms because the atomic mass of 35.45 is closer to the mass number of Cl-35. In fact, there are about three atoms of ${}^{35}_{17}Cl$ for every one atom of ${}^{37}_{17}Cl$ in a sample of chlorine atoms.

Table 3.8 lists the naturally occurring isotopes of some selected elements and their atomic masses.

TABLE 3.8 The Atomic Mass of Some Elements

Element	Isotopes	Atomic Mass (weighted average)	Most Prevalent Isotope
Lithium	${}^{6}_{3}Li$, ${}^{7}_{3}Li$	6.941 amu	${}^{7}_{3}Li$
Carbon	${}^{12}_{6}C$, ${}^{13}_{6}C$, ${}^{14}_{6}C$	12.01 amu	${}^{12}_{6}C$
Oxygen	${}^{16}_{8}O$, ${}^{17}_{8}O$, ${}^{18}_{8}O$	16.00 amu	${}^{16}_{8}O$
Fluorine	${}^{19}_{9}F$	19.00 amu	${}^{19}_{9}F$
Sulfur	${}^{32}_{16}S$, ${}^{33}_{16}S$, ${}^{34}_{16}S$, ${}^{36}_{16}S$	32.07 amu	${}^{32}_{16}S$
Copper	${}^{63}_{29}Cu$, ${}^{65}_{29}Cu$	63.55 amu	${}^{63}_{29}Cu$

CONCEPT CHECK 3.5

Average Atomic Mass

Neon consists of three naturally occurring isotopes: ${}^{20}_{10}Ne$, ${}^{21}_{10}Ne$, and ${}^{22}_{10}Ne$. Using the atomic mass on the periodic table, which isotope of neon is likely to be the most prevalent?

ANSWER

Using the periodic table, we find that the atomic mass for all the naturally occurring isotopes of neon is 20.18 amu. Since this number is very close to 20, the isotope Ne-20 that has a mass number of 20 is the most prevalent isotope in a naturally occurring sample of neon atoms.

SAMPLE PROBLEM 3.9

Calculating Atomic Mass

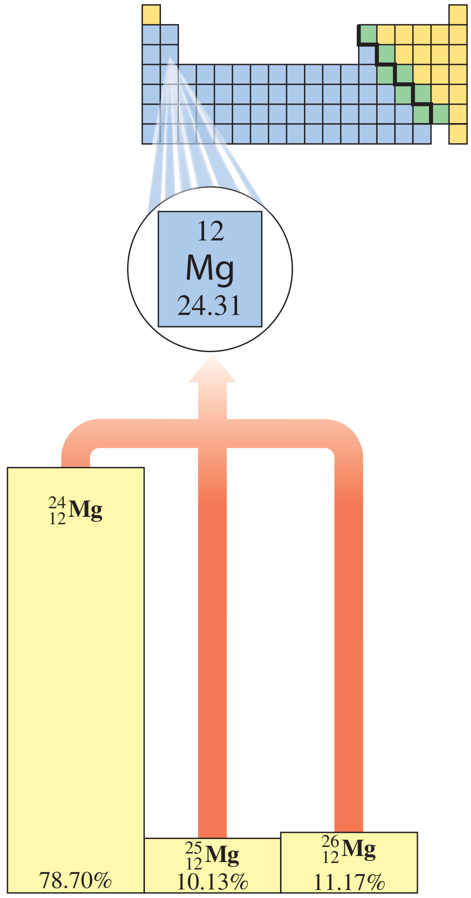
Using Table 3.7, calculate the atomic mass for magnesium using the weighted average mass method.

SOLUTION

Isotope	Mass (amu)		Abundance (%)		Contribution to the Atomic Mass
$^{24}_{12}Mg$	23.99	×	$\frac{78.70}{100}$	=	18.88 amu
$^{25}_{12}Mg$	24.99	×	$\frac{10.13}{100}$	=	2.531 amu
$^{26}_{12}Mg$	25.98	×	$\frac{11.17}{100}$	=	2.902 amu
	Atomic mass of Mg			=	24.31 amu (weighted average mass)

STUDY CHECK 3.9

There are two naturally occurring isotopes of boron. The isotope $^{10}_{5}B$ has a mass of 10.01 amu with an abundance of 19.80%, and the isotope $^{11}_{5}B$ has a mass of 11.01 amu with an abundance of 80.20%. What is the atomic mass of boron?



Magnesium, with three naturally occurring isotopes, has an atomic mass of 24.31 amu.

QUESTIONS AND PROBLEMS

Isotopes and Atomic Mass

3.31 What are the number of protons, neutrons, and electrons in the following isotopes?
a. $^{27}_{13}Al$ **b.** $^{52}_{24}Cr$ **c.** $^{34}_{16}S$ **d.** $^{81}_{35}Br$

3.32 What are the number of protons, neutrons, and electrons in the following isotopes?
a. $^{2}_{1}H$ **b.** $^{14}_{7}N$ **c.** $^{26}_{14}Si$ **d.** $^{70}_{30}Zn$

3.33 Write the atomic symbol for the isotope with each of the following characteristics:
a. 15 protons and 16 neutrons
b. 35 protons and 45 neutrons
c. 50 electrons and 72 neutrons
d. a chlorine atom with 18 neutrons
e. a mercury atom with 122 neutrons

3.34 Write the atomic symbol for the isotope with each of the following characteristics:
a. an oxygen atom with 10 neutrons
b. 4 protons and 5 neutrons
c. 25 electrons and 28 neutrons
d. a mass number of 24 and 13 neutrons
e. a nickel atom with 32 neutrons

3.35 There are three naturally occurring isotopes of argon, with mass numbers 36, 38, and 40.
a. Write the atomic symbol for each of these atoms.
b. How are these isotopes alike?
c. How are they different?
d. Why is the atomic mass of argon listed on the periodic table not a whole number?
e. Which isotope is the most prevalent in a sample of argon?

3.36 There are four isotopes of strontium with mass numbers 84, 86, 87, and 88.
a. Write the atomic symbol for each of these atoms.
b. How are these isotopes alike?
c. How are they different?
d. Why is the atomic mass of strontium listed on the periodic table not a whole number?
e. Which isotope is the most prevalent in a sample of strontium?

3.37 Two isotopes of gallium are naturally occurring, with $^{69}_{31}Ga$ at 60.11% (68.93 amu) and $^{71}_{31}Ga$ at 39.89% (70.92 amu). What is the atomic mass of gallium?

3.38 Two isotopes of copper are naturally occurring, with $^{63}_{29}Cu$ at 69.09% (62.93 amu) and $^{65}_{29}Cu$ at 30.91% (64.93 amu). What is the atomic mass of copper?

LEARNING GOAL

Given the name or symbol of one of the first 20 elements in the periodic table, write the electron arrangement.

TUTORIAL
Electromagnetic Radiation

TUTORIAL
Energy Levels

3.6 Electron Energy Levels

When we listen to a radio, use a microwave oven, turn on a light, see the colors of a rainbow, or have an X-ray, we are using various forms of *electromagnetic radiation.* Light and other electromagnetic radiation consist of energy particles that move as waves of energy. In a wave, just like the waves in an ocean, the distance between the peaks is called the *wavelength*. All forms of electromagnetic radiation travel in space at the speed of light, 3.0×10^8 m/s but differ in energy and wavelength. High-energy radiation has short wavelengths compared to low-energy radiation, which has longer wavelengths. The *electromagnetic spectrum* shows the arrangement of different types of electromagnetic radiation in order of increasing energy (see Figure 3.11).

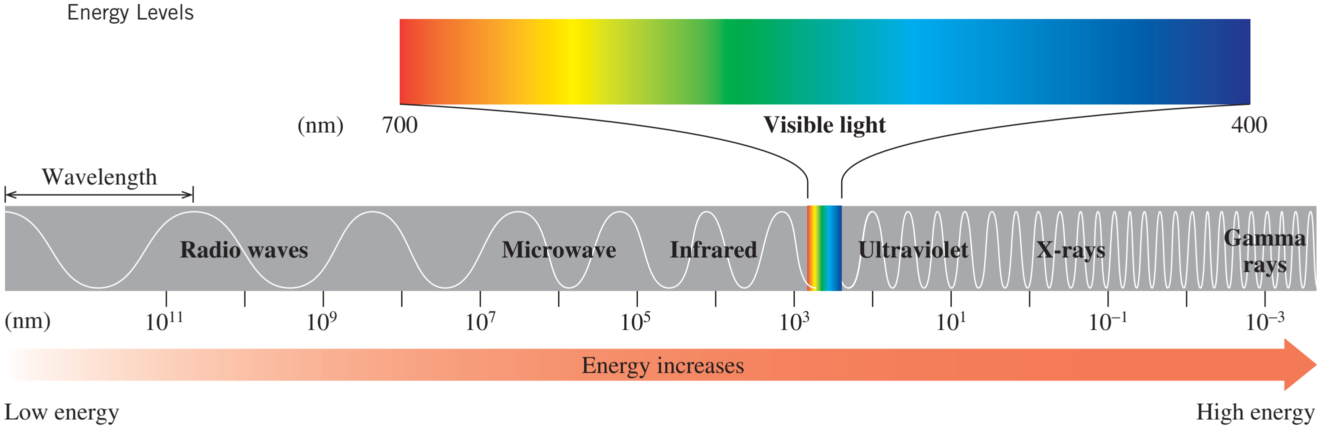
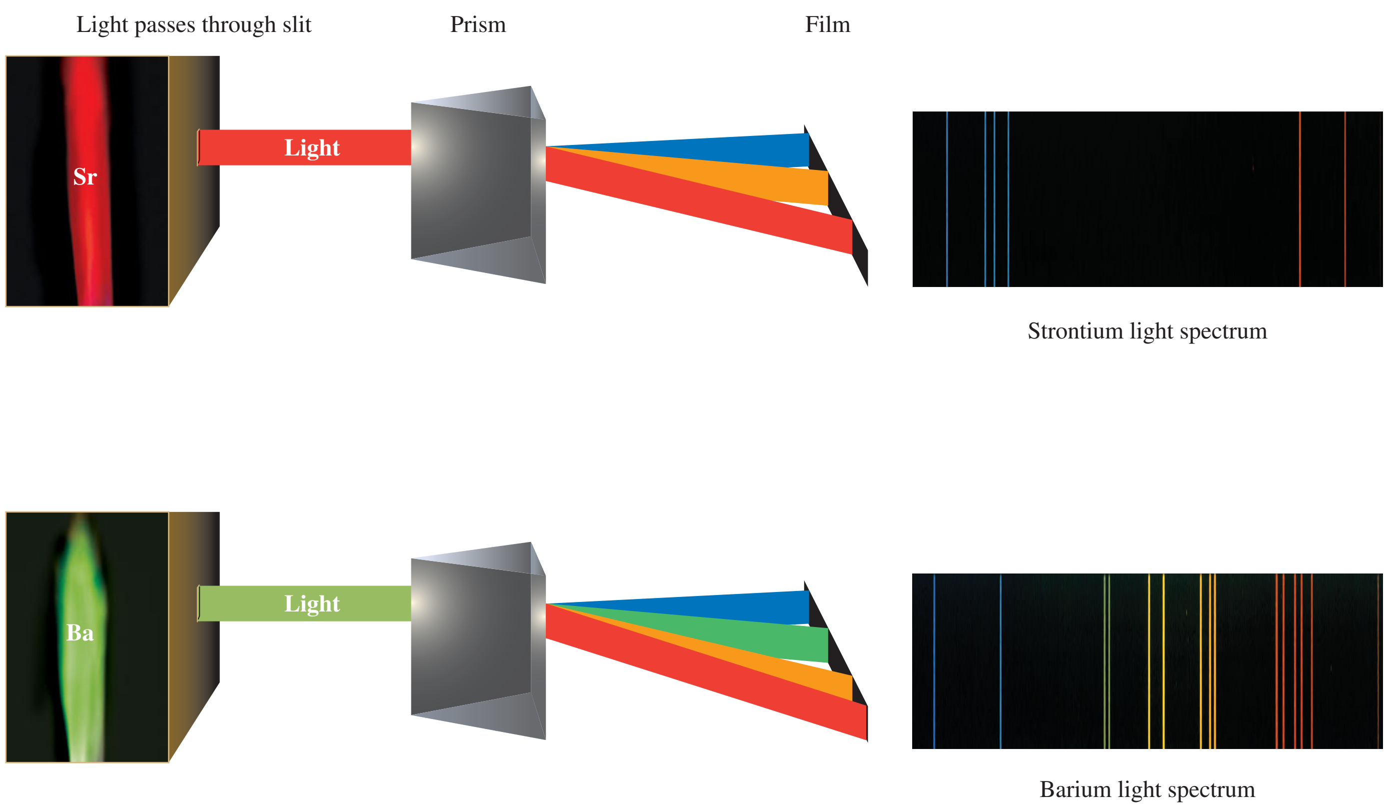


FIGURE 3.11 The electromagnetic spectrum shows the arrangement of wavelengths of electromagnetic radiation. The visible portion consists of wavelengths from 700 nm to 400 nm.

Q How does the wavelength of red light compare to that of blue light?



In an atomic spectrum, light from a heated element separates into distinct lines.

When the light from the Sun passes through a prism, the light separates into a continuous color spectrum, which consists of the colors we see in a rainbow. In contrast, when light from an element that is heated passes through a prism, it separates into distinct lines of color called an *atomic spectrum*. Each element has its own unique atomic spectrum.

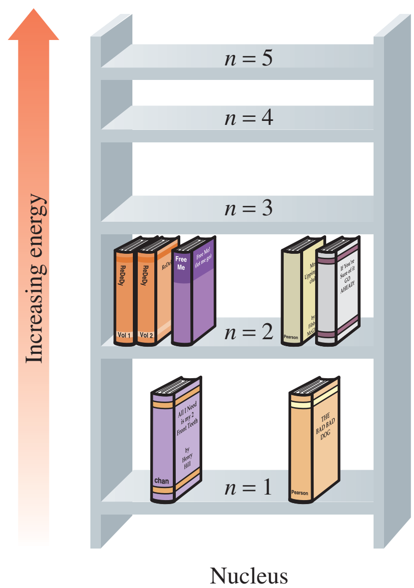
A rainbow forms when light passes through water droplets.

Electron Energy Levels

Scientists have now determined that the lines in the atomic spectra of elements are caused by changes in the energies of the electrons. In an atom, each electron has a specific energy known as its **energy level**. All the electrons with the same energy are grouped in the same energy level. The energy levels are assigned numbers (n) beginning with $n = 1$, $n = 2$, up to $n = 7$. Electrons in the lower energy levels are closer to the nucleus, whereas electrons in the higher energy levels are farther away.

As an analogy, we can think of the energy levels of an atom as similar to the shelves in a bookcase. The first shelf is the lowest energy level; the second shelf is the second energy level, and so on. If we are arranging books on the shelves, it would take less energy to fill the bottom shelf first, and then the second shelf, and so on. However, we could never get any book to stay in the space between any of the shelves. Similarly, the energy of an electron must be at specific energy levels, and not between.

Unlike bookcases, however, there is a large difference between the energy of the first and second levels, but then the higher energy levels are closer together. Another difference is that the lower electron energy levels hold fewer electrons than the higher energy levels.



An electron can have the energy of only one of the energy levels in an atom.

Chemistry Link to the Environment

ENERGY-SAVING FLUORESCENT BULBS

A compact fluorescent light bulb (CFL) is replacing the standard light bulb we use in our homes and workplaces. Compared to a standard light bulb, the CFL has a longer life and uses less electricity. Within about 20 days of use, the CFL saves enough money in electricity costs to pay for its higher initial cost.

A standard incandescent light bulb has a thin tungsten filament inside a sealed glass bulb. When the light is switched on, electricity flows through this filament, and electrical energy is converted to heat energy. When the filament reaches a temperature of about 2300 °C, we see white light.

A fluorescent bulb produces light in a different way. When the switch is turned on, electrons move between two electrodes and collide with mercury atoms in a mixture of mercury and argon gas inside the light. When the electrons in the mercury atoms absorb energy from the collisions, they are raised to higher energy levels. As electrons fall to lower energy levels, energy in the ultraviolet range is emitted. This ultraviolet light strikes the phosphor coating inside the tube, and fluorescence occurs as visible light is emitted.

The production of light in a fluorescent bulb is more efficient than in an incandescent light bulb. A 75-watt incandescent bulb can be replaced by a 20-watt CFL that gives the same amount of light, providing a 70% reduction in electricity costs. A typical incandescent light bulb lasts for one to two months, whereas a CFL lasts from one to two years. One drawback of the CFL is that each contains about 4 mg of mercury. As long as the bulb stays intact, no mercury is released. However, used CFL bulbs should not be disposed of in household trash but rather should be taken to a recycling center.

A compact fluorescent light bulb (CFL) uses up to 70% less energy.

Changes in Electron Energy Level

An electron can change from one energy level to a higher level only if it absorbs the energy equal to the difference between two levels. When an electron changes to a lower energy level, it emits energy equal to the difference between the two levels (see Figure 3.12). If the energy emitted is in the visible range, we see one of the colors of visible light. The yellow color of sodium streetlights and the red color of neon lights are examples of electrons emitting energy in the visible color range.

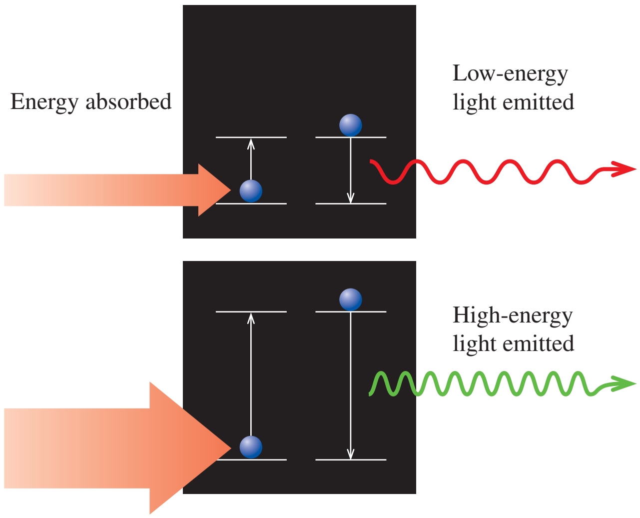


FIGURE 3.12 Electrons absorb a specific amount of energy to move to a higher energy level. When electrons lose energy, a specific quantity of energy is emitted.

Q What causes electrons to move to higher energy levels?

CONCEPT CHECK 3.6

Change in Energy Levels

a. How does an electron move to a higher energy level?

b. When an electron drops to a lower energy level, how is energy lost?

ANSWER

a. An electron moves to a higher energy level when it absorbs an amount of energy equal to the difference in energy levels.

b. Energy equal to the difference in energy levels is emitted when an electron drops to a lower energy level.

SELF STUDY ACTIVITY
Bohr's Shell Model of the Atom

Electron Arrangements for the First 20 Elements

The *electron arrangement* of an atom gives the number of electrons in each energy level. We can write the electron arrangements for the first 20 elements by placing electrons in energy levels beginning with the lowest. There is a limit to the number of electrons allowed in each energy level. Only a few electrons can occupy the lower energy levels, while more electrons can be accommodated in higher energy levels. We can now look at the numbers of electrons in the first four energy levels for the first 20 elements as shown in Table 3.9.

Period 1

H	1
He	2

The single electron of hydrogen goes into energy level 1, and the two electrons of helium fill energy level 1. Thus, energy level 1 can hold just two electrons. The electron arrangements for H and He are shown in the margin.

TABLE 3.9 Electron Arrangements for the First 20 Elements

			Number of Electrons in Energy Level			
Element	**Symbol**	**Atomic Number**	**1**	**2**	**3**	**4**
Hydrogen	H	1	1			
Helium	He	2	2			
Lithium	Li	3	2	1		
Beryllium	Be	4	2	2		
Boron	B	5	2	3		
Carbon	C	6	2	4		
Nitrogen	N	7	2	5		
Oxygen	O	8	2	6		
Fluorine	F	9	2	7		
Neon	Ne	10	2	8		
Sodium	Na	11	2	8	1	
Magnesium	Mg	12	2	8	2	
Aluminum	Al	13	2	8	3	
Silicon	Si	14	2	8	4	
Phosphorus	P	15	2	8	5	
Sulfur	S	16	2	8	6	
Chlorine	Cl	17	2	8	7	
Argon	Ar	18	2	8	8	
Potassium	K	19	2	8	8	1
Calcium	Ca	20	2	8	8	2

Period 2

For the elements of the second period (lithium, Li, to neon, Ne), we fill the first energy level with two electrons, and place the remaining electrons in the second energy level. For example, lithium has three electrons. Two of those electrons fill energy level 1. The remaining electron goes into the second energy level. We can write this electron arrangement as 2,1. Going across Period 2, more electrons are added to the second energy level. For example, an atom of carbon, with a total of six electrons, fills energy level 1, which leaves the four remaining electrons in the second energy level. The electron arrangement for carbon can be written 2,4. For neon, the last element in Period 2, both the first and second energy levels are filled to give an electron arrangement of 2,8.

Lithium	2,1
Carbon	2,4
Neon	2,8

Period 3

For sodium, the first element in Period 3, 10 electrons fill the first and second energy levels, which means that the remaining electron must go into the third energy level. The electron arrangement for sodium can be written 2,8,1. The elements that follow sodium in the third period continue to add electrons, one at a time, until the third level is complete. For example, a sulfur atom with 16 electrons fills the first and second energy levels, which leaves 6 electrons in the third level. The electron arrangement for sulfur can be written 2,8,6. At the end of Period 3, argon has 8 electrons in the third level. Once energy level 3 has 8 electrons, it stops filling. However, 10 more electrons will be added later.

Sodium	2,8,1
Sulfur	2,8,6
Argon	2,8,8

Period 4

For potassium, the first element in Period 4, electrons fill the first and second energy levels, and 8 electrons are in the third energy level. The remaining electron of potassium enters energy level 4, which gives the electron arrangement of 2,8,8,1. For calcium, the process is similar, except that calcium has 2 electrons in the fourth energy level. The electron arrangement for calcium can be written 2,8,8,2.

Potassium	2,8,8,1
Calcium	2,8,8,2

Elements Beyond 20

Following calcium, the next 10 electrons (scandium to zinc) are added to the third energy level, which already has 8 electrons, until it is complete with 18 electrons. The higher energy levels 5, 6, and 7 can theoretically accommodate up to 50, 72, and 98 electrons, but they are not completely filled. Beyond the first 20 elements, the electron arrangements become more complicated and will not be covered in this text.

Energy Level (n)	1	2	3	4	5	6	7
Number of Electrons	2	8	18	32	32	18	8

SAMPLE PROBLEM 3.10

Writing Electron Arrangements

Write the electron arrangement for each of the following:

a. oxygen **b.** chlorine

SOLUTION

a. Oxygen with atomic number 8, has eight electrons, which are arranged with two electrons in energy level 1 and six electrons in energy level 2.
2,6

b. Chlorine with atomic number 17, has 17 electrons, which are arranged with 2 electrons in energy level 1, 8 electrons in energy level 2, and 7 electrons in energy level 3.
2,8,7

STUDY CHECK 3.10

What element has an electron arrangement of 2,8,5?

QUESTIONS AND PROBLEMS

Electron Energy Levels

3.39 Electrons can move to higher energy levels when they ________ (absorb/emit) energy.

3.40 Electrons drop to lower energy levels when they ________ (absorb/emit) energy.

3.41 Identify the form of electromagnetic radiation in each pair that has the greater energy:
a. green light or yellow light
b. microwaves or blue light

3.42 Identify the form of electromagnetic radiation in each pair that has the greater energy:
a. radio waves or violet light
b. infrared light or ultraviolet light

3.43 Write the electron arrangement for each of the following elements:
(*Example*: sodium 2,8,1)
a. carbon **b.** argon **c.** potassium
d. silicon **e.** helium **f.** nitrogen

3.44 Write the electron arrangement for each of the following atoms:
(*Example*: sodium 2,8,1)
a. phosphorus **b.** neon **c.** sulfur
d. magnesium **e.** aluminum **f.** fluorine

3.45 Identify the elements that have the following electron arrangements:

	Energy Level			
	1	**2**	**3**	**4**
a.	2	1		
b.	2	8	2	
c.	1			
d.	2	8	7	
e.	2	6		

3.46 Identify the elements that have the following electron arrangements:

	Energy Level			
	1	**2**	**3**	**4**
a.	2	5		
b.	2	8	6	
c.	2	4		
d.	2	8	8	1
e.	2	8	3	

Chemistry Link to Health

BIOLOGICAL REACTIONS TO UV LIGHT

Our everyday life depends on sunlight, but exposure to sunlight can have damaging effects on living cells, and too much exposure can even cause their death. The light energy, especially ultraviolet (UV), excites electrons and may lead to unwanted chemical reactions. The list of damaging effects of sunlight includes sunburn; wrinkling; premature aging of the skin; changes in the DNA of the cells, which can lead to skin cancers; inflammation of the eyes; and perhaps cataracts. Some drugs, like the acne medications Accutane and Retin-A, as well as antibiotics, diuretics, sulfonamides, and estrogen, make the skin extremely sensitive to light.

However, medicine does take advantage of the beneficial effect of sunlight. Phototherapy can be used to treat certain skin conditions including psoriasis, eczema, and dermatitis. In the treatment of psoriasis, for example, oral drugs are given to make the skin more photosensitive; exposure to UV follows. Low-energy light is used to break down bilirubin in neonatal jaundice. Sunlight is also a factor in stimulating the immune system.

In cutaneous T-cell lymphoma, an abnormal increase in T cells causes painful ulceration of the skin. The skin is treated by photophoresis, in which the patient receives a photosensitive chemical, and then blood is removed from the body and exposed to ultraviolet light. The blood is returned to the patient, and the treated T cells stimulate the immune system to respond to the cancer cells.

In a disorder called seasonal affective disorder or SAD, people experience mood swings and depression during the winter. Some research suggests that SAD is the result of a decrease of serotonin, or an increase in melatonin, when there are fewer hours of sunlight. One treatment for SAD is therapy using bright light provided by a lamp called a light box. A daily exposure to intense light for 30 to 60 minutes seems to reduce symptoms of SAD.

3.7 Trends in Periodic Properties

LEARNING GOAL

Use the electron arrangement of elements to explain the trends in periodic properties.

The electron arrangements of atoms are an important factor in the physical and chemical properties of the elements. Now we will look at the *valence electrons* in atoms, *atomic size*, and *ionization energy*. Known as *periodic properties*, each increases or decreases across a period, and then the trend is repeated again in each successive period.

MC TUTORIAL Periodic Trends

Group Number and Valence Electrons

The chemical properties of representative elements in Groups 1A (1) to 8A (18) are mostly due to the **valence electrons**, which are the electrons in the outermost energy level. The **group number** gives the number of valence electrons for each group of representative elements. For example, all the elements in Group 1A (1) have one valence electron. All the elements in Group 2A (2) have two valence electrons. The halogens in Group 7A (17) all have seven valence electrons. Table 3.10 shows how the number of valence electrons for common representative elements is consistent with the group number.

CONCEPT CHECK 3.7

Using Group Numbers

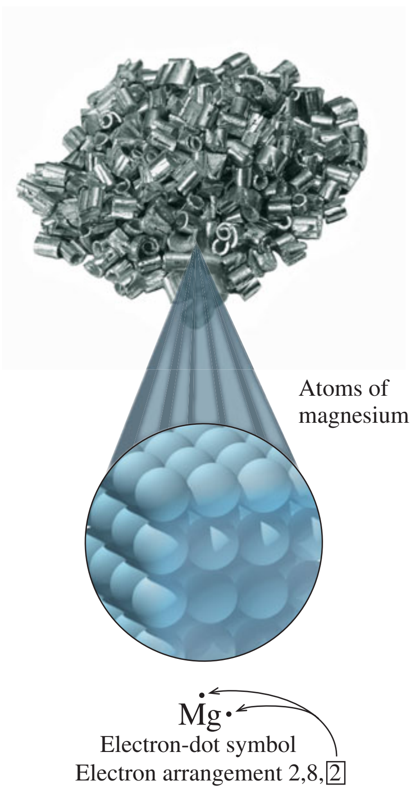
How can we determine that strontium has two valence electrons without writing its electron arrangement?

ANSWER

Strontium is in Group 2A (2). Because the group number is the same as the number of electrons in the outermost energy level, strontium would have two valence electrons.

TABLE 3.10 Comparison of Electron Arrangements, by Group, for Some Representative Elements

			Number of Electrons in Energy Level			
Group Number	**Element**	**Symbol**	**1**	**2**	**3**	**4**
1A (1)	Lithium	Li	2	**1**		
	Sodium	Na	2	8	**1**	
	Potassium	K	2	8	8	**1**
2A (2)	Beryllium	Be	2	**2**		
	Magnesium	Mg	2	8	**2**	
	Calcium	Ca	2	8	8	**2**
3A (13)	Boron	B	2	**3**		
	Aluminum	Al	2	8	**3**	
	Gallium	Ga	2	8	18	**3**
4A (14)	Carbon	C	2	**4**		
	Silicon	Si	2	8	**4**	
	Germanium	Ge	2	8	18	**4**
5A (15)	Nitrogen	N	2	**5**		
	Phosphorus	P	2	8	**5**	
	Arsenic	As	2	8	18	**5**
6A (16)	Oxygen	O	2	**6**		
	Sulfur	S	2	8	**6**	
	Selenium	Se	2	8	18	**6**
7A (17)	Fluorine	F	2	**7**		
	Chlorine	Cl	2	8	**7**	
	Bromine	Br	2	8	18	**7**
8A (18)	Helium	He	**2**			
	Neon	Ne	2	**8**		
	Argon	Ar	2	8	**8**	
	Krypton	Kr	2	8	18	**8**



SAMPLE PROBLEM 3.11

Using Group Numbers

Using the periodic table, write the group number and the number of valence electrons for each of the following elements:

a. cesium **b.** iodine

SOLUTION

a. Cesium (Cs) is in Group 1A (1); cesium has one valence electron.
b. Iodine (I) is in Group 7A (17); iodine has seven valence electrons.

STUDY CHECK 3.11

What is the group number of elements with atoms that have five valence electrons?

Electron-Dot Symbols

TUTORIAL
Electron-Dot Symbols for Elements

An **electron-dot symbol**, also known as a Lewis structure, represents the valence electrons as dots that are placed on the sides, top, or bottom of the symbol for the element. One to four valence electrons are arranged as single dots. When an atom has five to eight valence electrons, one or more electrons are paired. Any of the following would be an acceptable electron-dot symbol for magnesium, which has two valence electrons:

Possible Electron-Dot Symbols for the Two Valence Electrons in Magnesium

Ṁg· Ṁg ·Ṁg ·Mg· Mg· ·Mg

Electron-dot symbols for selected elements are given in Table 3.11.

Increasing number of valence electrons →

TABLE 3.11 Electron-Dot Symbols for Selected Elements in Periods 1–4

	Group Number							
	1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
Number of Valence Electrons	1	2	3	4	5	6	7	8*
Electron-Dot Symbol	H·							He:
	Li·	Be·	·Ḃ·	·Ċ·	·N̈·	·Ö:	·F̈:	:N̈e:
	Na·	Mg·	·Ȧl·	·Ṡi·	·P̈·	·S̈:	·C̈l:	:Ä̤r:
	K·	Ca·	·Ġa·	·Ġe·	·Äs·	·S̈e:	·B̈r:	:K̈r:

*Helium (He) is stable with 2 valence electrons.

Same number of valence electrons ↓

SAMPLE PROBLEM 3.12

Writing Electron-Dot Symbols

Write the electron-dot symbol for each of the following:

a. bromine **b.** aluminum

SOLUTION

a. Because the group number for bromine is 7A (17), bromine has seven valence electrons, which are drawn as seven dots, three pairs and one single dot, around the symbol Br.

·B̈r:

b. Aluminum, in Group 3A (13), has three valence electrons, which are shown as three single dots around the symbol Al.

·Ȧl·

STUDY CHECK 3.12

What is the electron-dot symbol for phosphorus?

Atomic Size

The size of an atom is determined by its *atomic radius*, which is the distance of the outermost electrons from the nucleus. For each group of representative elements, the atomic size *increases* from the top to the bottom because the outermost electrons in each energy level are farther from the nucleus. For example, in Group 1A (1), Li has a valence electron in energy level 2; Na has a valence electron in energy level 3; and K has a valence electron in energy level 4. This means that a K atom is larger than a Na atom, and a Na atom is larger than a Li atom (see Figure 3.13).

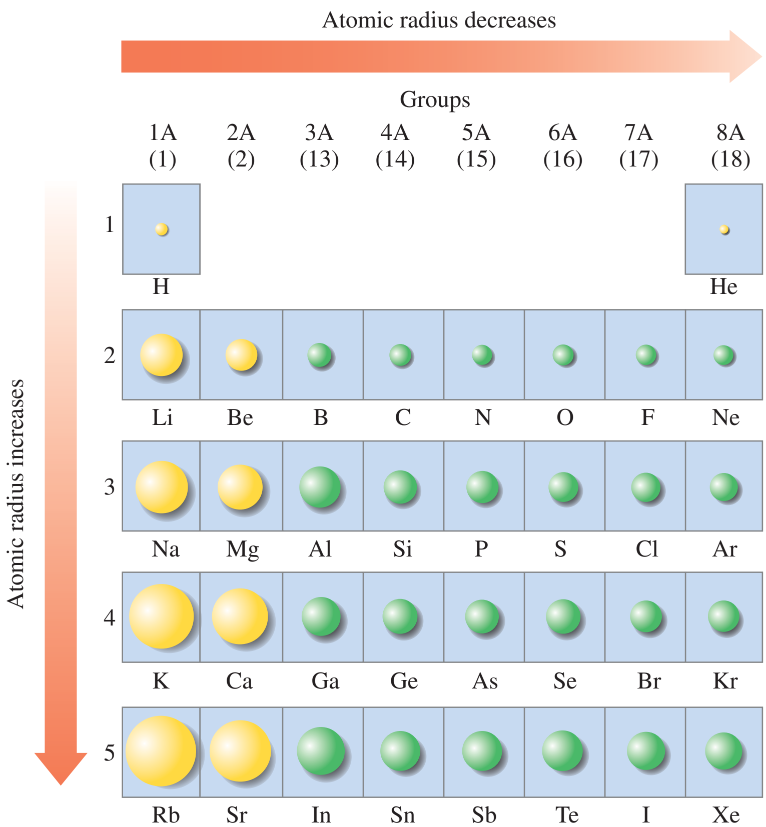


FIGURE 3.13 For representative elements, the atomic size increases going down a group but decreases going from left to right across a period.

Q Why does the atomic size increase going down a group of representative elements?

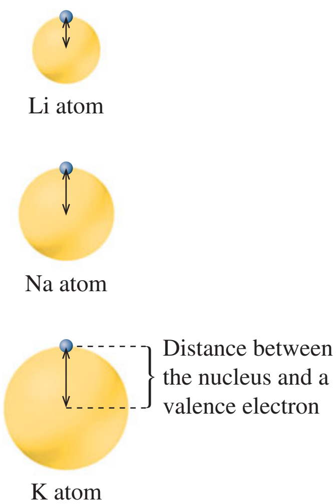
For the elements in a period, an increase in the number of protons in the nucleus increases the attraction for the outermost electrons. As a result, the outer electrons are pulled closer to the nucleus, which means that the size of representative elements decreases going from left to right across a period.

TUTORIAL
Ionization Energy

Ionization Energy

In an atom, negatively charged electrons are attracted to the positive charge of the protons in the nucleus. Thus, a quantity of energy known as the **ionization energy** is required to remove one of the outermost electrons. When an electron is removed from a neutral atom, a positive particle called a cation with a 1+ charge is formed.

$$Na(g) + \text{energy (ionization)} \longrightarrow Na^{+}(g) + e^{-}$$



As the distance from a valence electron to the nucleus increases in Group 1A (1), the ionization energy decreases.

The ionization energy *decreases* going down a group. Less energy is needed to remove an electron because nuclear attraction decreases when electrons are farther from the nucleus. Going across a period from left to right, the ionization energy *increases.* As the positive charge of the nucleus increases, more energy is needed to remove an electron.

In Period 1, the valence electrons are close to the nucleus and strongly held. H and He have high ionization energies because a large amount of energy is required to remove

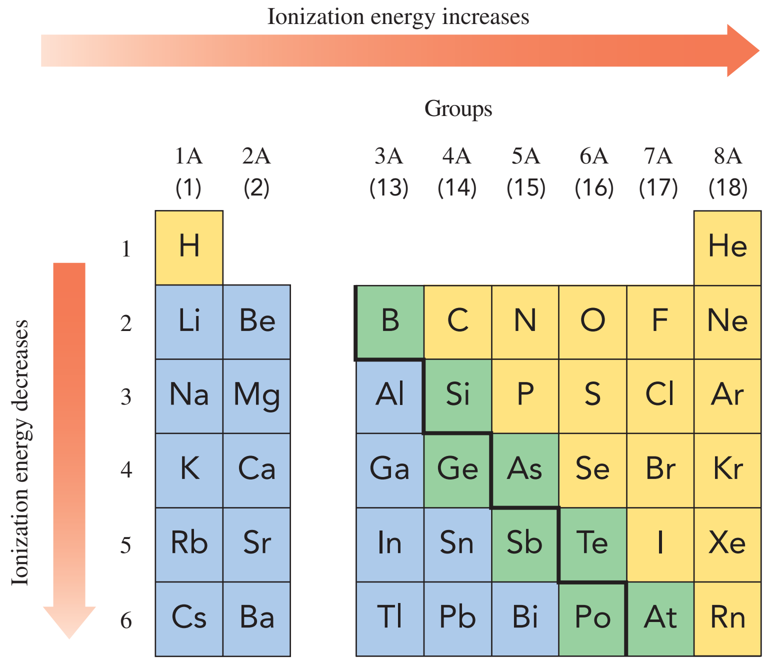


FIGURE 3.14 Ionization energies for the representative elements tend to decrease going down a group and increase going across a period.

Q Why is the ionization energy for F greater than for Cl?

an electron. The ionization energy for He is the highest of any element because He has a full, stable, energy level which is disrupted by removing an electron. The high ionization energies of the noble gases indicate that their electron arrangements are especially stable. In general, the ionization energy is low for the metals and high for the nonmetals (see Figure 3.14).

SAMPLE PROBLEM 3.13

Ionization Energy

Indicate the element in each group that has the higher ionization energy and explain your choice.

a. K or Na **b.** Mg or Cl **c.** F, N, or C

SOLUTION

a. Na. In Na, the valence electron is closer to the nucleus.
b. Cl. Attraction for the valence electrons increases going from left to right across a period.
c. F. Attraction for the valence electrons increases going from left to right across a period.

STUDY CHECK 3.13

Arrange Sn, Sr, and I in order of increasing ionization energy.

Metallic Character

In Section 3.2, we identified elements as metals, nonmetals, and metalloids. An element that has **metallic character** is an element that loses valence electrons easily. Metallic character is more prevalent in the elements (metals) on the left side of the period table and decreases going from the left side to the right side of the periodic table. The elements (nonmetals) on the right side of the periodic table do not easily lose electrons,

TUTORIAL
Patterns in the Periodic Table

which means they are the least metallic. Most of the metalloids between the metals and nonmetals tend to lose electrons, but not as easily as the metals. Thus, in Period 3, sodium, which loses electrons most easily, would be the most metallic. Going across from left to right in Period 3, metallic character decreases to argon, which has the least metallic character.

For elements in the same group of representative elements, metallic character increases going from top to bottom. Atoms at the bottom of any group have more electron levels, which makes it easier to lose electrons. Thus, the elements at the bottom of a group on the periodic table have lower ionization energy and are more metallic compared to the elements at the top (see Figure 3.15).

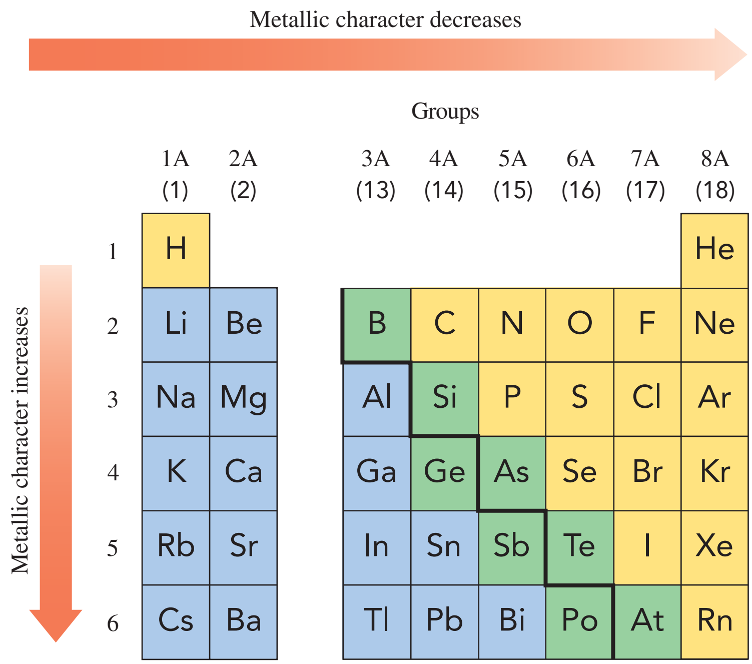


FIGURE 3.15 Metallic character of the representative elements increases going down a group and decreases going from left to right across a period.

Q Why is the metallic character greater for Rb than for Li?

A summary of the trends in periodic properties we have discussed is given in Table 3.12.

TABLE 3.12 Summary of Trends in Periodic Properties of Representative Elements

Periodic Property	Top to Bottom of a Group	Left to Right across a Period
Valence Electrons	Remains the same	Increases
Atomic Radius	Increases due to the increase in number of energy levels	Decreases due to the increase of protons in the nucleus that pull electrons closer
Ionization Energy	Decreases because outer electrons are easier to remove when they are farther away from the nucleus	Increases as the attraction of the protons for outer electrons requires more energy to remove an electron
Metallic Character	Increases because outer electrons are easier to remove when they are farther away from the nucleus	Decreases as the attraction of the protons makes it more difficult to remove an electron

CONCEPT CHECK 3.8

Metallic Character

Identify the element that has more metallic character in each of the following:

a. Mg or Al **b.** Na or K

ANSWER

a. Mg is more metallic than Al because metallic character decreases from left to right across a period.

b. K is more metallic than Na because metallic character increases going down a group.

QUESTIONS AND PROBLEMS

Trends in Periodic Properties

3.47 What is the group number and number of valence electrons for each of the following elements?
a. magnesium **b.** chlorine **c.** oxygen
d. nitrogen **e.** barium **f.** bromine

3.48 What is the group number and number of valence electrons for each of the following elements?
a. lithium **b.** silicon **c.** neon
d. argon **e.** tin **f.** boron

3.49 Write the group number and electron-dot symbol for each of the following elements:
a. sulfur **b.** nitrogen **c.** calcium
d. sodium **e.** gallium

3.50 Write the group number and electron-dot symbol for each of the following elements:
a. carbon **b.** oxygen **c.** argon
d. lithium **e.** chlorine

3.51 Place the elements in each set in order of decreasing atomic radius.
a. Al, Si, Mg **b.** Cl, I, Br **c.** Sr, Sb, I **d.** P, Si, Na

3.52 Place the elements in each set in order of decreasing atomic radius.
a. Cl, S, P **b.** Ge, Si, C **c.** Ba, Ca, Sr **d.** S, O, Se

3.53 Select the larger atom in each pair.
a. Na or Cl **b.** Na or Rb **c.** Na or Mg **d.** Rb or I

3.54 Select the larger atom in each pair.
a. S or Cl **b.** S or O **c.** S or Se **d.** S or Mg

3.55 Arrange each set of elements in order of increasing ionization energy.
a. F, Cl, Br **b.** Na, Cl, Al **c.** Na, K, Cs **d.** As, Ca, Br

3.56 Arrange each set of elements in order of increasing ionization energy.
a. O, N, C **b.** S, P, Cl **c.** As, P, N **d.** Al, Si, P

3.57 Select the element in each pair with the higher ionization energy.
a. Br or I **b.** Mg or Sr **c.** Si or P **d.** I or Xe

3.58 Select the element in each pair with the higher ionization energy.
a. O or Ne **b.** K or Br **c.** Ca or Ba **d.** N or O

3.59 Fill in the following blanks using larger or smaller, more metallic or less metallic. Na has a ________ atomic size and is ________ than P.

3.60 Fill in the following blanks using larger or smaller, lower or higher. Mg has a ________ atomic size and a ________ ionization energy than Cs.

3.61 Place the following in order of decreasing metallic character:
Br, Ge, Ca, Ga

3.62 Place the following in order of increasing metallic character:
Mg, P, Al, Ar

3.63 Fill in each of the following blanks using higher or lower, more or less:
Sr has a ________ ionization energy and ________ metallic character than Sb.

3.64 Fill in each of the following blanks using higher or lower, more or less:
N has a ________ ionization energy and ________ metallic character than As.

3.65 Complete each of the statements **a** through **d** using 1, 2, or 3:
1. decreases **2.** increases **3.** remains the same
Going down Group 6A (16),
a. the ionization energy ________
b. the atomic size ________
c. the metallic character ________
d. the number of valence electrons ________

3.66 Complete each of the statements **a** through **d** using 1, 2, or 3:
1. decreases **2.** increases **3.** remains the same
Going from left to right across Period 4,
a. the ionization energy ________
b. the atomic size ________
c. the metallic character ________
d. the number of valence electrons ________

3.67 Which statements completed with **a** through **e** will be true?

In Period 2, the nonmetals compared to the metals have larger (greater)

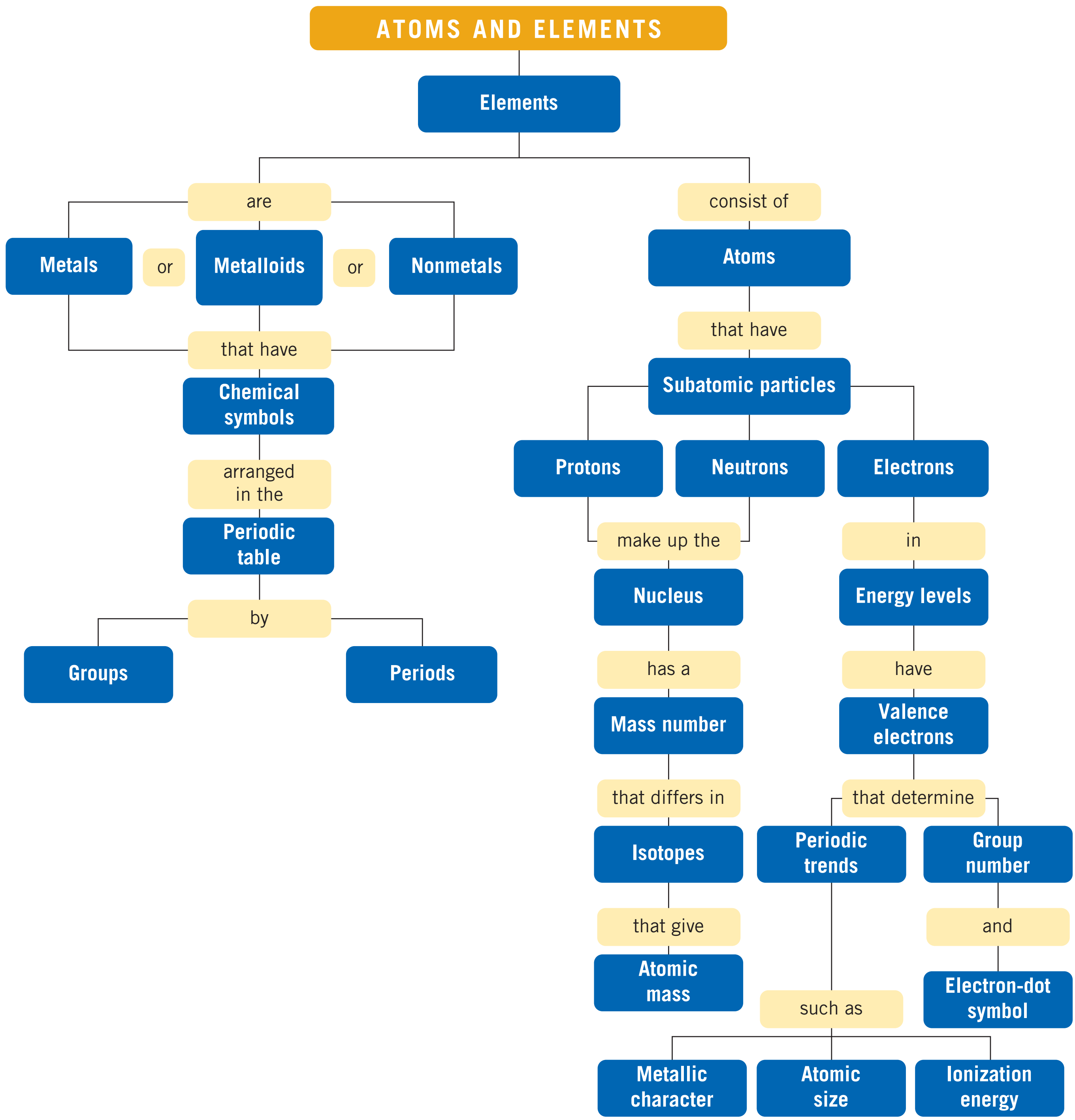
a. atomic size
b. ionization energies
c. number of protons
d. metallic character
e. number of valence electrons

3.68 Which statements completed with **a** through **e** will be true?

In Group 4A (14), an atom of C compared to an atom of Sn has a larger (greater)

a. atomic size
b. ionization energy
c. number of protons
d. metallic character
e. number of valence electrons

CONCEPT MAP



CHAPTER REVIEW

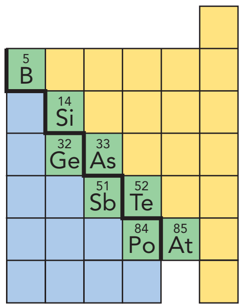
3.1 Elements and Symbols

Learning Goal: Given the name of an element, write its correct symbol; from the symbol, write the correct name.

Elements are the primary substances of matter. Chemical symbols are one- or two-letter abbreviations of the names of the elements.

3.2 The Periodic Table

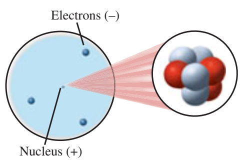
Learning Goal: Use the periodic table to identify the group and the period of an element; identify the element as a metal, nonmetal, or metalloid.



The periodic table is an arrangement of the elements by increasing atomic number. A vertical column or group on the periodic table contains elements with similar properties. A horizontal row is called a period. Elements in Group 1A (1) are called the alkali metals; Group 2A (2), alkaline earth metals; Group 7A (17), the halogens; and Group 8A (18), the noble gases. On the periodic table, metals are located on the left of the heavy zigzag line, and nonmetals are to the right of the heavy zigzag line. Except for aluminum, elements located on the heavy line are called metalloids.

3.3 The Atom

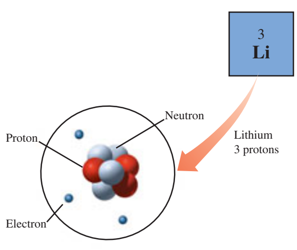
Learning Goal: Describe the electrical charge and location in an atom for a proton, a neutron, and an electron.



An atom is the smallest particle that retains the characteristics of an element. Within atoms, there are three subatomic particles. Protons have a positive charge (+), electrons carry a negative charge (−), and neutrons are electrically neutral. The protons and neutrons are found in the tiny, dense nucleus whereas electrons are located in a large space surrounding the nucleus.

3.4 Atomic Number and Mass Number

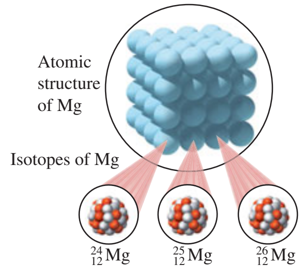
Learning Goal: Given the atomic number and the mass number of an atom, state the number of protons, neutrons, and electrons.



The atomic number gives the number of protons in all the atoms of the same element. In a neutral atom, the number of protons and the number of electrons are equal. The mass number is the total number of protons and neutrons in an atom.

3.5 Isotopes and Atomic Mass

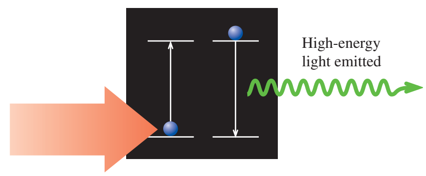
Learning Goal: Give the number of protons, electrons, and neutrons in an isotope of an element; calculate the atomic mass of an element using the abundance and mass of its naturally occurring isotopes.



Atoms that have the same number of protons, but different numbers of neutrons, are called isotopes. The atomic mass of an element is the weighted average mass of all the isotopes in a naturally occurring sample of that element.

3.6 Electron Energy Levels

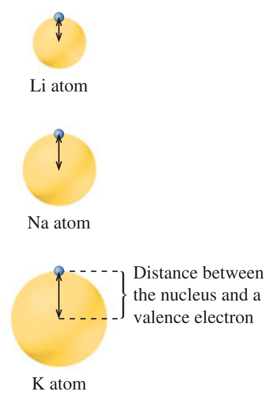
Learning Goal: Given the name or symbol of one of the first 20 elements in the periodic table, write the electron arrangement.



Every electron has a specific amount of energy. In an atom, the electrons of similar energy are grouped in specific energy levels. The first level nearest the nucleus can hold 2 electrons, the second level can hold 8 electrons, and the third level will take up to 18 electrons. The electron arrangement is written by placing the number of electrons in that atom in order from the lowest energy levels and filling to higher levels.

3.7 Trends in Periodic Properties

Learning Goal: Use the electron arrangement of elements to explain the trends in periodic properties.



The similarity of behavior for the elements in a group is related to having the same number of valence electrons, which are the electrons in the outermost energy level. The group number for an element gives the number of valence electrons, which means that each group of elements has the same arrangement of valence electrons differing only in the energy level. The size of an atom increases going down a group and decreases going across a period. The energy required to remove a valence electron is the ionization energy, which decreases going down a group and generally increases going from left to right across a period. The metallic character increases going down a group and decreases going across a period.

Key Terms

alkali metals Elements of Group 1A (1) except hydrogen; these are soft, shiny metals with one outer shell electron.

alkaline earth metals Group 2A (2) elements, which have 2 electrons in their outer shells.

atom The smallest particle of an element that retains the characteristics of the element.

atomic mass The weighted average mass of all the naturally occurring isotopes of an element.

atomic mass unit (amu) A small mass unit used to describe the mass of very small particles such as atoms and subatomic particles; 1 amu is equal to one-twelfth the mass of a ${}^{12}_{6}C$ atom.

atomic number A number that is equal to the number of protons in an atom.

atomic symbol An abbreviation used to indicate the mass number and atomic number of an isotope.

chemical symbol An abbreviation that represents the name of an element.

electron A negatively charged subatomic particle having a very small mass that is usually ignored in calculations; its symbol is e^-.

electron-dot symbol The representation of an atom that shows valence electrons as dots around the symbol of the element.

energy level A group of electrons with similar energy.

group A vertical column in the periodic table that contains elements having similar physical and chemical properties.

group number A number that appears at the top of each vertical column (group) in the periodic table and indicates the number of electrons in the outermost energy level.

halogens Group 7A (17) elements fluorine, chlorine, bromine, iodine, and astatine.

ionization energy The energy needed to remove the least tightly bound electron from the outermost energy level of an atom.

isotope An atom that differs only in mass number from another atom of the same element. Isotopes have the same atomic number (number of protons), but different numbers of neutrons.

mass number The total number of neutrons and protons in the nucleus of an atom.

metal An element that is shiny, malleable, ductile, and a good conductor of heat and electricity. The metals are located to the left of the zigzag line in the periodic table.

metallic character A measure of how easily an element loses a valence electron.

metalloid Elements with properties of both metals and nonmetals, located along the zigzag line on the periodic table.

neutron A neutral subatomic particle having a mass of about 1 amu and found in the nucleus of an atom; its symbol is n or n^0.

noble gas An element in Group 8A (18) of the periodic table, generally unreactive and seldom found in combination with other elements.

nonmetal An element with little or no luster that is a poor conductor of heat and electricity. The nonmetals are located to the right of the zigzag line in the periodic table.

nucleus The compact, very dense center of an atom, containing the protons and neutrons of the atom.

period A horizontal row of elements in the periodic table.

periodic table An arrangement of elements by increasing atomic number such that elements having similar chemical behavior are grouped in vertical columns.

proton A positively charged subatomic particle having a mass of about 1 amu and found in the nucleus of an atom; its symbol is p or p^+.

representative elements Elements found in Groups 1A (1) through 8A (18) excluding B groups (3–12) of the periodic table.

transition elements Elements located between Groups 2A (2) and 3A (13) on the periodic table.

valence electrons Electrons in the outermost energy level of an atom.

Understanding the Concepts

3.69 According to Dalton's atomic theory, which of the following are true?
a. Atoms of an element are identical to atoms of other elements.
b. Every element is made of atoms.
c. Atoms of two different elements combine to form compounds.
d. In a chemical reaction, some atoms disappear and new atoms appear.

3.70 Use Rutherford's gold-foil experiment to answer each of the following:
a. What did Rutherford expect to happen when he aimed particles at the gold foil?
b. How did the results differ from what he expected?
c. How did he use the results to propose a model of the atom?

3.71 Use the subatomic particles (**1**–**3**) to define each of the following:
1. protons **2.** neutrons **3.** electrons
a. atomic mass
b. atomic number
c. positive charge
d. negative charge
e. mass number – atomic number

3.72 Use the subatomic particles (**1**–**3**) to define each of the following:
1. protons **2.** neutrons **3.** electrons
a. mass number
b. surround the nucleus
c. in the nucleus
d. charge of 0
e. equal to number of electrons

3.73 Consider the following atoms in which X represents the chemical symbol of the element:

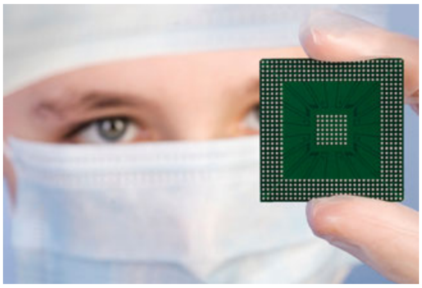
$${}^{16}_{8}X \quad {}^{16}_{9}X \quad {}^{18}_{10}X \quad {}^{17}_{8}X \quad {}^{18}_{8}X$$

a. What atoms have the same number of protons?
b. Which atoms are isotopes? Of what element?
c. Which atoms have the same mass number?
d. What atoms have the same number of neutrons?

3.74 Cadmium, atomic number 48, consists of eight naturally occurring isotopes. Do you expect any of the isotopes to have the atomic mass listed on the periodic table for cadmium? Explain.

3.75 Indicate if the atoms in each pair have the same number of protons? Neutrons? Electrons?
a. ${}^{37}_{17}Cl$, ${}^{38}_{18}Ar$ **b.** ${}^{36}_{14}Si$, ${}^{35}_{14}Si$ **c.** ${}^{40}_{18}Ar$, ${}^{39}_{17}Cl$

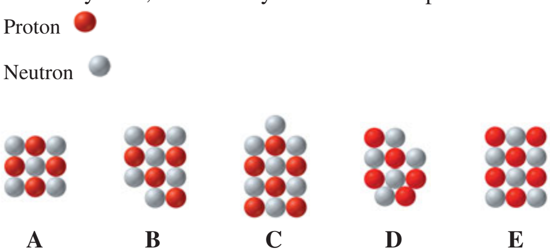
3.76 Complete the following table for the three naturally occurring isotopes of silicon, the major component in computer chips:



Computer chips consist primarily of the element silicon.

	Isotope		
	$^{28}_{14}Si$	$^{29}_{14}Si$	$^{30}_{14}Si$
Number of protons			
Number of neutrons			
Number of electrons			
Atomic number			
Mass number			

3.77 For each representation of a nucleus **A** through **E**, write the atomic symbol, and identify which are isotopes.



3.78 Identify the element represented by each nucleus **A** through **E** in Problem 3.77 as a metal, nonmetal, or metalloid.

3.79 Match the spheres **A** through **D** with atoms of Li, Na, K, and Rb.

3.80 Match the spheres **A** through **D** with atoms of K, Ge, Ca, and Kr.

3.81 Of the elements Na, Mg, Si, S, Cl, and Ar, identify one that fits each of the following:
a. largest atomic size
b. a halogen
c. electron arrangement 2,8,4
d. highest ionization energy
e. in Group 6A (16)
f. most metallic character
g. two valence electrons

3.82 Of the elements Sn, Xe, Te, Sr, I, and Rb, identify one that fits each of the following:
a. smallest atomic size
b. an alkaline earth metal
c. a metalloid
d. lowest ionization energy
e. in Group 4A (14)
f. least metallic character
g. seven valence electrons

Additional Questions and Problems

For instructor-assigned homework, go to **www.masteringchemistry.com**.

3.83 Give the group number and period number for each of the following elements:
a. bromine **b.** argon **c.** potassium **d.** radium

3.84 Give the group number and period number for each of the following elements:
a. radon **b.** arsenic **c.** carbon **d.** neon

3.85 Indicate if each of the following statements is true or false:
a. The proton is a negatively charged particle.
b. The neutron is 2000 times as heavy as a proton.
c. The atomic mass unit is based on a carbon atom with 6 protons and 6 neutrons.
d. The nucleus is the largest part of the atom.
e. The electrons are located outside the nucleus.

3.86 Indicate if each of the following statements is true or false:
a. The neutron is electrically neutral.
b. Most of the mass of an atom is due to the protons and neutrons.
c. The charge of an electron is equal, but opposite, to the charge of a neutron.
d. The proton and the electron have about the same mass.
e. The mass number is the number of protons.

3.87 Complete the following statements:
a. The atomic number gives the number of ________ in the nucleus.
b. In an atom, the number of electrons is equal to the number of ________.
c. Sodium and potassium are examples of elements called ________.

3.88 Complete the following statements:
a. The number of protons and neutrons in an atom is also the ________ number.
b. The elements in Group 7A (17) are called the ________.
c. Elements that are shiny and conduct heat are called ________.

3.89 Write the names and symbols of the elements with the following atomic numbers:
a. 28 **b.** 56 **c.** 88 **d.** 33
e. 50 **f.** 55 **g.** 79 **h.** 80

3.90 Write the names and symbols of the elements with the following atomic numbers:
a. 10 **b.** 22 **c.** 48 **d.** 26
e. 54 **f.** 78 **g.** 83 **h.** 92

3.91 For the following atoms, give the number of protons, neutrons, and electrons:

a. $^{107}_{47}Ag$ **b.** $^{98}_{43}Tc$ **c.** $^{208}_{82}Pb$ **d.** $^{222}_{86}Rn$ **e.** $^{136}_{54}Xe$

3.92 For the following atoms, give the number of protons, neutrons, and electrons:

a. $^{198}_{79}Au$ **b.** $^{127}_{53}I$ **c.** $^{75}_{35}Br$ **d.** $^{133}_{55}Cs$ **e.** $^{195}_{78}Pt$

3.93 Complete the following table:

Name	Atomic Symbol	Number of Protons	Number of Neutrons	Number of Electrons
	$^{34}_{16}S$			
		28	34	
Magnesium			14	
	$^{220}_{86}Rn$			

3.94 Complete the following table:

Name	Atomic Symbol	Number of Protons	Number of Neutrons	Number of Electrons
Potassium			22	
	$^{51}_{23}V$			
		48	64	
Barium			82	

3.95 Write the atomic symbol for each of the following:
a. an atom with 4 protons and 5 neutrons
b. an atom with 12 protons and 14 neutrons
c. a calcium atom with a mass number of 46
d. an atom with 30 electrons and 40 neutrons
e. a copper atom with 34 neutrons

3.96 Write the atomic symbol for each of the following:
a. an aluminum atom with 14 neutrons
b. an atom with atomic number 26 and 32 neutrons
c. a strontium atom with 50 neutrons
d. an atom with a mass number of 72 and atomic number 33

3.97 The most abundant isotope of lead is $^{208}_{82}Pb$.
a. How many protons, neutrons, and electrons are in $^{208}_{82}Pb$?
b. What is the atomic symbol of another isotope of lead with 132 neutrons?
c. What is the atomic symbol and name of an atom with the same mass number as in part **b** and 131 neutrons?

3.98 The most abundant isotope of silver is $^{107}_{47}Ag$.
a. How many protons, neutrons, and electrons are in $^{107}_{47}Ag$?
b. What is the atomic symbol of another isotope of silver with 62 neutrons?
c. What is the atomic symbol and name of an atom with the same mass number as in part **b** and 61 neutrons?

3.99 Write the group number and the electron arrangement for each of the following:
a. oxygen **b.** sodium
c. neon **d.** boron

3.100 Write the group number and the electron arrangement for each of the following:
a. magnesium **b.** chlorine
c. beryllium **d.** argon

3.101 Why is the ionization energy of Ca higher than K, but lower than Mg?

3.102 Why is the ionization energy of Cl lower than F, but higher than S?

3.103 Of the elements Li, Be, N, and F, which
a. is an alkaline earth metal?
b. has the largest atomic radius?
c. has the highest ionization energy?
d. is found in Group 5A (15)?
e. has the most metallic character?

3.104 Of the elements F, Br, Cl, I, which
a. is the largest atom?
b. is the smallest atom?
c. has the lowest ionization energy?
d. requires the most energy to remove an electron?
e. is found in Period 4?

Challenge Questions

3.105 For each of the following, write the symbol and name for X and the number of protons and neutrons. Which are isotopes of each other?

a. $^{37}_{17}X$ **b.** $^{56}_{27}X$ **c.** $^{116}_{50}X$ **d.** $^{124}_{50}X$ **e.** $^{116}_{48}X$

3.106 There are four naturally occurring isotopes of iron: $^{54}_{26}Fe$, $^{56}_{26}Fe$, $^{57}_{26}Fe$, and $^{58}_{26}Fe$.
a. How many protons, neutrons, and electrons are in $^{58}_{26}Fe$?
b. What is the most prevalent isotope in an iron sample?
c. How many neutrons are in $^{57}_{26}Fe$?
d. Why don't any of the isotopes of iron have the atomic mass of 55.85 amu listed on the periodic table?

3.107 Give the symbol of the element that has the
a. smallest atomic size in Group 6A (16)
b. smallest atomic size in Period 3
c. highest ionization energy in Group 5A (15)
d. lowest ionization energy in Period 3
e. most metallic character in Group 2A (2)

3.108 Give the symbol of the element that has the
a. largest atomic size in Group 1A (1)
b. largest atomic size in Period 4
c. highest ionization energy in Group 2A (2)
d. lowest ionization energy in Group 7A (17)
e. least metallic character in Group 4A (14)

3.109 A lead atom has a mass of 3.4×10^{-22} g. How many lead atoms are in a cube of lead that has a volume of 2.00 cm^3 if the density of lead is 11.3 g/cm^3?

3.110 If the diameter of a sodium atom is 3.14×10^{-8} cm, how many sodium atoms would fit along a line exactly 1 inch long?

3.111 Silicon has three naturally occurring isotopes: Si-28 (27.977 amu) with an abundance of 92.23%, Si-29 (28.976 amu) with a 4.68% abundance, and Si-30 (29.974 amu) with a 3.09% abundance. What is the atomic mass of silicon?

3.112 Antimony (Sb), which has an atomic weight of 121.75 amu, has two naturally occurring isotopes: Sb-121 and Sb-123. If a sample of antimony is 42.70% Sb-123, which has a mass of 122.90 amu, what is the mass of Sb-121?

Answers

Answers to Study Checks

3.1 Si, S, and Ag

3.2 magnesium, aluminum, and fluorine

3.3 **a.** Strontium is in Group 2A (2).
b. Strontium is an alkaline earth metal.
c. Strontium is in Period 5.
d. Magnesium, Mg
e. Alkali metal, Rb; halogen, I; noble gas, Xe

3.4 **a.** arsenic, As **b.** argon, Ar **c.** silicon, Si

3.5 False; the nucleus occupies a very small volume in an atom.

3.6 **a.** 79 **b.** 79 **c.** gold, Au

3.7 45 neutrons

3.8 **a.** $^{15}_{7}N$ **b.** $^{42}_{20}Ca$ **c.** $^{27}_{13}Al$

3.9 10.81 amu

3.10 phosphorus

3.11 Group 5A (15)

3.12 $\cdot\ddot{\underset{\cdot}{P}}\cdot$

3.13 Ionization energy increases going left to right across a period: Sr is lowest, Sn is higher, and I is the highest of this set.

Answers to Selected Questions and Problems

3.1 **a.** Cu **b.** Pt **c.** Ca **d.** Mn
e. Fe **f.** Ba **g.** Pb **h.** Sr

3.3 **a.** carbon **b.** chlorine **c.** iodine **d.** mercury
e. silver **f.** argon **g.** boron **h.** nickel

3.5 **a.** sodium, chlorine
b. calcium, sulfur, oxygen
c. carbon, hydrogen, chlorine, nitrogen, oxygen
d. calcium, carbon, oxygen

3.7 **a.** Period 2 **b.** Group 8A (18)
c. Group 1A (1) **d.** Period 2

3.9 **a.** alkaline earth metal **b.** transition element
c. noble gas **d.** alkali metal
e. halogen

3.11 **a.** C **b.** He **c.** Na **d.** Ca **e.** Al

3.13 **a.** metal **b.** nonmetal **c.** metal **d.** nonmetal
e. nonmetal **f.** nonmetal **g.** metalloid **h.** metal

3.15 **a.** electron **b.** proton **c.** electron **d.** neutron

3.17 Rutherford determined that an atom contains a small, compact nucleus that is positively charged.

3.19 **a.** True **b.** True
c. True **d.** False; a proton is attracted to an electron

3.21 In the process of brushing hair, strands of hair become charged with like charges that repel each other.

3.23 **a.** atomic number **b.** both
c. mass number **d.** atomic number

3.25 **a.** lithium, Li **b.** fluorine, F
c. calcium, Ca **d.** zinc, Zn
e. neon, Ne **f.** silicon, Si
g. iodine, I **h.** oxygen, O

3.27 **a.** 18 protons and 18 electrons
b. 30 protons and 30 electrons
c. 53 protons and 53 electrons
d. 19 protons and 19 electrons

3.29 See Table 3.13.

3.31 **a.** 13 protons, 14 neutrons, 13 electrons
b. 24 protons, 28 neutrons, 24 electrons
c. 16 protons, 18 neutrons, 16 electrons
d. 35 protons, 46 neutrons, 35 electrons

3.33 **a.** $^{31}_{15}P$ **b.** $^{80}_{35}Br$ **c.** $^{122}_{50}Sn$
d. $^{35}_{17}Cl$ **e.** $^{202}_{80}Hg$

3.35 **a.** $^{36}_{18}Ar$ $^{38}_{18}Ar$ $^{40}_{18}Ar$
b. They all have the same number of protons and electrons.
c. They have different numbers of neutrons, which gives them different mass numbers.
d. The atomic mass of Ar listed on the periodic table is the average atomic mass of all the isotopes.
e. Because argon has an atomic mass of 39.95, the isotope of $^{40}_{18}Ar$ would be the most prevalent.

3.37 69.72 amu

3.39 absorb

3.41 **a.** green light **b.** blue light

3.43 **a.** 2,4 **b.** 2,8,8 **c.** 2,8,8,1
d. 2,8,4 **e.** 2 **f.** 2,5

3.45 **a.** Li **b.** Mg **c.** H
d. Cl **e.** O

TABLE 3.13

Name of the Element	Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
Aluminum	Al	13	27	13	14	13
Magnesium	Mg	12	24	12	12	12
Potassium	K	19	39	19	20	19
Sulfur	S	16	31	16	15	16
Iron	Fe	26	56	26	30	26

3.47 **a.** Group 2A (2), 2 e^-
b. Group 7A (17), 7 e^-
c. Group 6A (16), 6 e^-
d. Group 5A (15), 5 e^-
e. Group 2A (2), 2 e^-
f. Group 7A (17), 7 e^-

3.49 **a.** Group 6A (16)
$\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{S}}\cdot$
b. Group 5A (15)
$\cdot\overset{\cdot\cdot}{\underset{\cdot}{N}}\cdot$
c. Group 2A (2)
$\cdot Ca \cdot$
d. Group 1A (1)
$Na\cdot$
e. Group 3A (13)
$\cdot\overset{\cdot}{Ga}\cdot$

3.51 **a.** Mg, Al, Si **b.** I, Br, Cl **c.** Sr, Sb, I **d.** Na, Si, P

3.53 **a.** Na **b.** Rb **c.** Na **d.** Rb

3.55 **a.** Br, Cl, F **b.** Na, Al, Cl **c.** Cs, K, Na **d.** Ca, As, Br

3.57 **a.** Br **b.** Mg **c.** P **d.** Xe

3.59 larger, more metallic

3.61 Ca, Ga, Ge, Br

3.63 lower, more

3.65 **a.** decreases **b.** increases
c. increases **d.** remains the same

3.67 **b**, **c**, and **e** are true.

3.69 **b** and **c** are true.

3.71 **a.** 1 + 2 **b.** 1 **c.** 1 **d.** 3 **e.** 2

3.73 **a.** $^{16}_{8}X$, $^{17}_{8}X$, $^{18}_{8}X$ all have 8 protons.
b. $^{16}_{8}X$, $^{17}_{8}X$, $^{18}_{8}X$ are all isotopes of O.
c. $^{16}_{8}X$ and $^{16}_{9}X$ have mass numbers of 16, and $^{18}_{10}X$ and $^{18}_{8}X$ have mass numbers of 18.
d. $^{16}_{8}X$ and $^{18}_{10}X$ have 8 neutrons each.

3.75 **a.** Both have 20 neutrons.
b. Both have 14 protons and 14 electrons.
c. Both have 22 neutrons.

3.77 **a.** $^{9}_{4}Be$ **b.** $^{11}_{5}B$ **c.** $^{13}_{6}C$ **d.** $^{10}_{5}B$ **e.** $^{12}_{6}C$
Representations **B** and **D** are isotopes of boron; **C** and **E** are isotopes of carbon.

3.79 Li is **D**, Na is **A**, K is **C**, and Rb is **B**.

3.81 **a.** Na **b.** Cl **c.** Si **d.** Ar
e. S **f.** Na **g.** Mg

3.83 **a.** Group 7A (17), Period 4
b. Group 8A (18), Period 3
c. Group 1A (1), Period 4
d. Group 2A (2), Period 7

3.85 **a.** False **b.** False **c.** True
d. False **e.** True

3.87 **a.** protons **b.** protons **c.** alkali metals

3.89 **a.** nickel, Ni **b.** barium, Ba
c. radium, Ra **d.** arsenic, As
e. tin, Sn **f.** cesium, Cs
g. gold, Au **h.** mercury, Hg

3.91 **a.** 47 protons, 60 neutrons, 47 electrons
b. 43 protons, 55 neutrons, 43 electrons
c. 82 protons, 126 neutrons, 82 electrons
d. 86 protons, 136 neutrons, 86 electrons
e. 54 protons, 82 neutrons, 54 electrons

3.93

Name	Atomic Symbol	Number of Protons	Number of Neutrons	Number of Electrons
Sulfur	$^{34}_{16}S$	16	18	16
Nickel	$^{62}_{28}Ni$	28	34	28
Magnesium	$^{26}_{12}Mg$	12	14	12
Radon	$^{220}_{86}Rn$	86	134	86

3.95 **a.** $^{9}_{4}Be$ **b.** $^{26}_{12}Mg$ **c.** $^{46}_{20}Ca$
d. $^{70}_{30}Zn$ **e.** $^{63}_{29}Cu$

3.97 **a.** 82 protons, 126 neutrons, 82 electrons
b. $^{214}_{82}Pb$
c. $^{214}_{83}Bi$, bismuth

3.99 **a.** O; Group 6A (16); 2,6
b. Na; Group 1A (1); 2,8,1
c. Ne; Group 8A (18); 2,8
d. B; Group 3A (13); 2,3

3.101 Calcium has a greater number of protons than K. The increase in positive charge increases the attraction for electrons, which means that more energy is required to remove the valence electrons. The valence electrons are farther from the nucleus in Ca than in Mg and less energy is needed to remove them.

3.103 **a.** Be **b.** Li **c.** F
d. N **e.** Li

3.105 **a.** Cl, chlorine; 17 protons and 20 neutrons
b. Co, cobalt; 27 protons and 29 neutrons
c. Sn, tin; 50 protons and 66 neutrons
d. Sn, tin; 50 protons and 74 neutrons
e. Cd, cadmium; 48 protons and 68 neutrons
The symbols **c** and **d** are isotopes of Sn, tin.

3.107 **a.** O **b.** Ar **c.** N
d. Na **e.** Ra

3.109 6.6×10^{22} atoms of Pb

3.111 28.09 amu