CENGAGE

TENTH EDITION

CHEMISTRY FOR TODAY

GENERAL, ORGANIC, AND BIOCHEMISTRY

SEAGER | SLABAUGH | HANSEN

TENTH EDITION

Chemistry for Today

General, Organic, and Biochemistry

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***Chemistry for Today: General, Organic, and Biochemistry,* 10th Edition**
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This 10th edition is dedicated to the memory of Wayne Aprill, a high school science teacher who touched the lives of everyone around him; students, family and many good friends.

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Maren S. Hansen has been a science teacher at West High School, where she taught honors biology. She has also taught AP biology and biology in the International Baccalaureate Program. She received her B.A. and master of education degrees from Weber State University. Her professional interests have focused upon helping students participate in Science Olympiad and Science Fair. Other interests include adventure travel, mountain hiking, gardening, and the company of friends and family. She hopes to share her love of science with her two children. She currently serves as an adjunct professor at the University of South Dakota, where she teaches online courses in general chemistry, elementary organic chemistry, and elementary biochemistry.

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Preface

The Image of Chemistry

We, as authors, are pleased that the acceptance of the previous nine editions of this textbook by students and their teachers has made it possible to publish this tenth edition. In the earlier editions, we expressed our concern about the negative image of chemistry held by many of our students, and their genuine fear of working with chemicals in the laboratory. Unfortunately, this negative image not only persists, but seems to be intensifying. Reports in the media related to chemicals or to chemistry continue to be primarily negative, and in many cases seem to be designed to increase the fear and concern of the general public. With this edition, we continue to hope that those who use this book will gain a more positive understanding and appreciation of the important contributions that chemistry makes in their lives.

Theme and Organization

This edition continues the theme of the positive and useful contributions made by chemistry in our world.

This text is designed to be used in either a two-semester or three-quarter course of study that provides an introduction to general chemistry, organic chemistry, and biochemistry. Most students who take such courses are majoring in nursing, other health professions, or the life sciences, and consider biochemistry to be the most relevant part of the course of study. However, an understanding of biochemistry depends upon a sound background in organic chemistry, which in turn depends upon a good foundation in general chemistry. We have attempted to present the general and organic chemistry in sufficient depth and breadth to make the biochemistry understandable.

The decisions about what to include and what to omit from the text were based on our combined 75-plus years of teaching, input from numerous reviewers and adopters, and our philosophy that a textbook functions as a personal tutor to each student. In the role of a personal tutor, a text must be more than just a collection of facts, data, and exercises. It should also help students relate to the material they are studying, carefully guide them through more difficult material, provide them with interesting and relevant examples of chemistry in their lives, and become a reference and a resource that they can use in other courses or their professions.

New to This Edition

In this tenth edition of the text, we have some exciting new features, including Career Focus boxes written by Monica Linford and Health Connections. We have also retained features that received a positive reception from our own students, the students of other adopters, other teachers, and reviewers. Over 133 figures and 170 examples have been added to the chapters. Many of these new figures and examples are health-related. Moreover, 559 end-of-chapter exercises have been added.

Also new to this edition are many new photographs and updated art to further enhance student comprehension of key concepts, processes, and test preparation.

Revision Summary of Tenth Edition

Chapter 1

- New Career Focus
- New Career Description
- 8 new or revised figures
- 11 new or revised Examples
- New photography
- 22 new Exercises

Chapter 2

- New Career Focus
- New Career Description
- 5 new or revised Examples
- 7 new or revised figures
- New photography
- 10 new Exercises

Chapter 3

- New Career Focus
- New Career Description
- 4 new or revised Examples
- 4 new or revised figures
- New photography
- 10 revised Exercises

Chapter 4

- New Career Focus
- New Career Description
- 10 new or revised Examples
- 6 new or revised figures
- New photography
- 12 new or revised Exercises

Chapter 5

- New Career Focus
- New Career Description
- 4 new or revised figures
- 6 new or revised Examples
- New photography
- 11 new or revised Exercises

Chapter 6

- New Career Focus
- New Career Description
- 8 new or revised Examples
- 8 new or revised figures
- New photography
- 22 new Exercises

Chapter 7

- New Career Focus
- New Career Description
- 8 new or revised figures
- New photography
- 7 new or revised Examples
- 9 new Exercises

Chapter 8

- New Career Focus
- New Career Description
- 3 new or revised Examples
- Several revised figures
- New photography

Chapter 9

- New Career Focus
- New Career Description
- 7 new or revised Examples
- 12 new or revised figures
- New photography
- 19 new Exercises

Chapter 10

- New Career Focus
- New Career Description
- 3 new or revised Examples
- 7 new or revised figures
- New photography
- 5 new Exercises

Chapter 11

- New Career Focus
- New Career Description
- 9 new or revised Examples
- 9 new or revised figures
- New photography
- 30 new Exercises

Chapter 12

- New Career Focus
- New Career Description
- 7 new or revised Examples
- 7 new or revised figures
- New photography
- 29 new Exercises

Chapter 13

- New Career Focus
- New Career Description
- 12 new or revised Examples
- 6 new or revised figures
- New photography
- 33 new Exercises

Chapter 14

- New Career Focus
- New Career Description
- 6 new or revised Examples
- New photography
- 36 new Exercises

Chapter 15

- New Career Focus
- New Career Description
- 4 new or revised Examples
- 5 new or revised figures
- New photography
- 31 new Exercises

Chapter 16

- New Career Focus
- New Career Description
- 11 new or revised Examples
- 5 new or revised figures
- New photography
- 31 new Exercises

Chapter 17

- New Career Focus
- New Career Description
- 7 new or revised Examples
- 4 new or revised figures
- New photography
- 30 new Exercises

Chapter 18

- New Career Focus
- New Career Description
- 10 new or revised Examples
- 7 new or revised figures
- New photography
- 30 new Exercises

Chapter 19

- New Career Focus
- New Career Description
- 7 new or revised Examples
- 5 new or revised figures
- New photography
- 30 new Exercises

Chapter 20

- New Career Focus
- New Career Description
- 5 new or revised Examples
- 6 new or revised figures
- New photography
- 30 new Exercises

Chapter 21

- New Career Focus
- New Career Description
- 5 new or revised Examples
- New photography
- 30 new Exercises

Chapter 22

- New Career Focus
- New Career Description
- 6 new or revised Examples
- Figure revised to show a current food label
- New photography
- 30 new Exercises

Chapter 23

- New Career Focus
- New Career Description
- 5 new or revised Examples
- New photography
- 30 new Exercises

Chapter 24

- New Career Focus
- New Career Description
- 8 new or revised Examples
- New photography
- 30 new Exercises

Chapter 25

- New Career Focus
- New Career Description
- 4 new or revised Examples
- New photography
- 30 new Exercises

Features

Each chapter has features especially designed to help students study effectively, as well as organize, understand, and enjoy the material in the course.

Career Focus. These features introduce the diverse fields of health care. The purpose of the career focus features is to stimulate inquiry; for that reason, we've placed them at the beginning of each chapter of the book. Vocabulary and scenarios may be unfamiliar to you who are studying these course materials, but our intent is to raise interest and pique your curiosity. A career description can be found at the end of each chapter before the Concept Summary.

Health Career Focus

STERILE PROCESSING TECHNICIAN

"Without us, nothing happens in this hospital. We're essential to every treatment, procedure, and surgery that occurs in this facility."

"Here's my locker where I change into my PPE, or personal protective equipment—just as if I were dressing for surgery—including foot coverings and a face shield."

"There, behind that window, we receive trays of used surgical equipment. We meticulously wash the blood, bone, and tissue off and get them ready for steam-pressure sterilization. This job isn't for the squeamish."

"One of the most important parts of my job is actually counting the instruments, making sure each item returns from surgery and is recorded carefully in the database. I also monitor equipment for

Chapter Outlines and Learning Objectives. At the beginning of each chapter, a list of learning objectives provides students with a convenient overview of what they should gain by studying the chapter. In order to help students navigate through each chapter and focus on key concepts, these objectives are repeated at the beginning of the section in which the applicable information is discussed. The objectives are referred to again in question format the concept summary at the end of each chapter. Thus, students begin each chapter with a set of objectives and end with an indication of how well they satisfied the objectives.

Key Terms. Identified within the text by the use of bold type, key terms are defined in the margin near the place where they are introduced. Students reviewing a chapter can quickly identify the important concepts on each page with this marginal glossary. A full glossary of key terms and concepts appears at the end of the text.

Environmental Connections. These boxed features contain current chemistry-related environmental issues such as "Ozone: Good Up High, Bad Nearby" and "CO_2 Emissions: A Blanket around the Earth."

Health Connections. These boxed features contain current chemistry-related health issues such as "Add Color to Your Diet," and suggestions for maintaining good health such as "Consider the Mediterranean Diet," "Cut Back on Processed Meat," and "Try a Little Chocolate."

HEALTH CONNECTIONS 14.1

Faking a Tan

Many people believe that a suntan makes one look healthy and attractive. Studies, however, indicate that this perception is far from the truth. According to these studies, sunbathing, especially when sunburn results, ages the skin prematurely and increases the risk of skin cancer. Cosmetic companies have developed a tanning alternative for those not willing to risk using the sun but who want to be "fashionably" tan.

Tanning lotions and creams that chemically darken the skin are now available. The active ingredient in these "bronzers" is dihydroxyacetone (DHA), a colorless compound classified by the Food and Drug Administration as a safe skin dye.

$$HO-CH_2-\overset{\overset{\displaystyle O}{\|}}{C}-CH_2-OH$$

dihydroxyacetone (DHA)

Within several hours after application, DHA produces a brown skin color by reacting with the outer layer of the skin, which consists of dead cells. Only the dead cells react with DHA, so the color gradually fades as the dead cells slough off and are replaced. This process generally leads to the fading of chemical tans within a few weeks. Another problem with chemical tans is uneven skin color. Areas of skin such as elbows and knees, which contain a thicker layer of dead cells, may absorb and react with more tanning lotion and become darker than other areas.

Perhaps the greatest problem with chemical tans is the false sense of security they might give. Some people with chemical tans think it is safe to go into the sun and get a deeper tan. This isn't true. Sunlight presents the same hazards to chemically tanned skin that it does to untanned skin.

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Some DHA-containing products.

Examples. To reinforce students in their problem-solving skill development, complete step-by-step solutions for numerous examples are included.

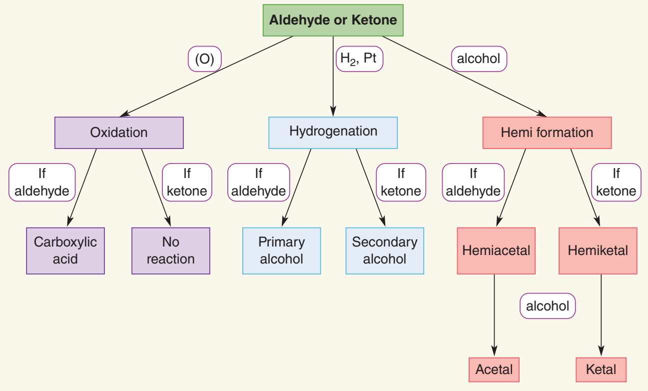
Learning Checks. Short self-check exercises follow examples and discussions of key or difficult concepts. A complete set of solutions is included in Appendix C. These allow students to measure immediately their understanding and progress.

Study Tools. Most chapters contain a *Study Tools* feature in which a challenging topic, skill, or concept of the chapter is addressed. Study suggestions, analogies, and approaches are provided to help students master these ideas.

STUDY TOOLS 14.1

A Reaction Map for Aldehydes and Ketones

This reaction map is designed to help you master organic reactions. Whenever you are trying to complete an organic reaction, use these two basic steps: (1) Identify the functional group that is to react, and (2) Identify the reagent that is to react with the functional group. If the reacting functional group is an aldehyde or a ketone, find the reagent in the summary diagram, and use the diagram to predict the correct products.



Concept Summary. Located at the end of each chapter, this feature provides a concise review of the concepts and challenges students to check their achievement of the learning objectives related to the concepts.

Concept Summary

14.1 The Nomenclature of Aldehydes and Ketones

Learning Objective: Can you recognize the carbonyl group in compounds and classify the compounds as aldehydes or ketones? Can you assign IUPAC name to aldehydes and ketones?

- The functional groups characteristic of aldehydes and ketones are very similar.

$$-\overset{\overset{\displaystyle O}{\|}}{C}-$$

- They both contain a carbonyl group.
- Aldehydes have a hydrogen attached to the carbonyl carbon.

$$-\overset{\overset{\displaystyle O}{\|}}{C}-H$$

- The polarity of the carbonyl group and the fact that aldehydes and ketones can form hydrogen bonds with water explain why the low-molecular-weight compounds of these organic classes are water-soluble.

14.3 Chemical Properties

Learning Objective: Can you write key reactions for aldehydes and ketones?

- Aldehydes are prepared by the oxidation of primary alcohols.
- Ketones are prepared by the oxidation of secondary alcohols.
- Aldehydes can be further oxidized to carboxylic acids, but ketones resist oxidation.
- Thus, aldehydes are oxidized by Tollens' reagent (Ag^+) and Benedict's solution (Cu^{2+}), whereas ketones are not.
- A characteristic reaction of both aldehydes and ketones is the addition of hydrogen to the carbonyl double bond to form alcohols.

Key Terms and Concepts. These are listed at the end of each chapter for easy review, with a reference to the chapter section in which they are presented.

Key Equations. This feature provides a useful summary of general equations and reactions from the chapter. This feature is particularly helpful to students in the organic chemistry chapters.

Exercises. Nearly 2200 end-of-chapter exercises are arranged by section. Approximately half of the exercises are answered in the back of the text. Solutions and answers to all exercises are provided in the Instructor's Manual. We have included a significant number of clinical and other familiar applications of chemistry in the exercises.

Chemistry for Thought. Included at the end of each chapter are special questions designed to encourage students to expand their reasoning skills. Some of these exercises are based on photographs found in the chapter, while others emphasize clinical or other useful applications of chemistry.

Possible Course Outlines

This text may be used effectively in either a two-semester or three-quarter course of study:

First semester: Chapters 1–13 (general chemistry and three chapters of organic chemistry)

Second semester: Chapters 14–25 (organic chemistry and biochemistry)

First semester: Chapters 1–10 (general chemistry)

Second semester: Chapters 11–21 (organic chemistry and some biochemistry)

First quarter: Chapters 1–10 (general chemistry)

Second quarter: Chapters 11–18 (organic chemistry)

Third quarter: Chapters 19–25 (biochemistry)

Supporting Materials

Additional instructor resources for this product are available online. Instructor assets include an Instructor's Manual including solutions to exercises, an Educator's Guide describing digital homework assets, a Transition Guide (9th to 10th edition), PowerPoint® slides, and a Test Bank powered by Cognero®. Sign up or sign in at **www.cengage.com** to search for and access this product and its online resources.

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Spencer L. Seager
Michael R. Slabaugh
Maren S. Hansen

1 Matter, Measurements, and Calculations

VanoVasaio/Shutterstock.com

Health Career Focus

STERILE PROCESSING TECHNICIAN

"Without us, nothing happens in this hospital. We're essential to every treatment, procedure, and surgery that occurs in this facility."

"Here's my locker where I change into my PPE, or personal protective equipment—just as if I were dressing for surgery—including foot coverings and a face shield."

"There, behind that window, we receive trays of used surgical equipment. We meticulously wash the blood, bone, and tissue off and get them ready for steam-pressure sterilization. This job isn't for the squeamish."

"One of the most important parts of my job is actually counting the instruments, making sure each item returns from surgery and is recorded carefully in the database. I also monitor equipment for proper temperatures and settings to ensure we meet protocols."

"This job requires extreme attention to detail. I initial everything I sterilize. My work is traceable to me. Precision is important."

Follow-up to this Career Focus appears at the end of the chapter before the *Concept Summary*.

Learning Objectives

When you complete the study of this chapter, you should be able to:

1. Explain what matter is. **(Section 1.1)**
2. Explain the difference between the terms *physical* and *chemical* as applied to the properties of matter and changes in matter. **(Section 1.2)**
3. Describe matter in terms of the accepted scientific model. **(Section 1.3)**
4. On the basis of observation or information given to you, classify matter into the correct category of each of the following pairs: heterogeneous or homogeneous, solution or pure substance, and element or compound. **(Section 1.4)**
5. Recognize the use of measurement units in everyday activities. **(Section 1.5)**

6 Recognize units of the metric system, and convert measurements done using the metric system into related units. **(Section 1.6)**

7 Express numbers using scientific notation and do calculations with numbers expressed in scientific notation. **(Section 1.7)**

8 Express the results of measurements and calculations using the correct number of significant figures. **(Section 1.8)**

9 Use the factor-unit method to solve numerical problems. **(Section 1.9)**

10 Do calculations involving percentages. **(Section 1.10)**

11 Do calculations involving densities. **(Section 1.11)**

CHEMISTRY is often described as the scientific study of matter. In a way, almost every study is a study of matter, because matter is the substance of everything. However, chemists are especially interested in matter; they study it and attempt to understand it from nearly every possible point of view.

The chemical nature of all matter makes an understanding of chemistry useful and necessary for individuals who are studying in a wide variety of areas, including the health sciences, natural sciences, home economics, education, environmental science, and law enforcement.

Matter comes in many shapes, sizes, and colors that are interesting to look at and describe. Early chemists did little more than describe what they observed, and their chemistry was a descriptive science that was severely limited in scope. It became a much more useful science when chemists began to make quantitative measurements, do calculations, and incorporate the results into their descriptions. Some fundamental ideas about matter are presented in this chapter, along with some ideas about quantitative measurement, the scientific measurement system, and calculations.

1.1 What Is Matter?

Learning Objective 1 Explain what matter is.

Definitions are useful in all areas of knowledge. They provide a common vocabulary for both presentations to students and discussions between professionals. You will be expected to learn a number of definitions as you study chemistry, and the first one is a definition of *matter*. Earlier, we said that matter is the substance of everything. That isn't very scientific, even though we think we know what it means. If you stop reading for a moment and look around, you will see a number of objects that might include people, potted plants, walls, furniture, books, windows, and a TV set or radio. The objects you see have at least two things in common: Each one has mass, and each one occupies space. These two common characteristics provide the basis for the scientific definition of matter. **Matter** is anything that has mass and occupies space. You probably understand what is meant by an object occupying space, especially if you have tried to occupy

matter Anything that has mass and occupies space.

the same space as some other object. The resulting physical bruises leave a lasting mental impression.

You might not understand the meaning of the term *mass* quite as well, but it can also be illustrated "painfully." Imagine walking into a very dimly lit room and being able to just barely see two large objects of equal size on the floor. You know that one is a bowling ball and the other is an inflated plastic ball, but you can't visually identify which is which. However, a hard kick delivered to either object easily allows you to identify each one. The bowling ball resists being moved much more strongly than does the inflated ball. Resistance to movement depends on the amount of matter in an object, and **mass** is an actual measurement of the amount of matter present.

mass A measurement of the amount of matter in an object.

The term *weight* is probably more familiar to you than *mass*, but the two are related. All objects are attracted to each other by gravity, and the greater their mass, the stronger the attraction between them. The **weight** of an object on Earth is a measurement of the gravitational force pulling the object toward Earth. An object with twice the mass of a second object is attracted with twice the force, and therefore has a weight twice the weight of the second object. The mass of an object is constant, no matter where it is located (even if it is in a weightless condition in outer space). However, the weight of an object depends on the strength of the gravitational attraction to which it is subjected. For example, a rock that weighs 16 pounds on Earth would weigh about 2.7 pounds on the moon (see **Figure 1.1**) because the gravitational attraction on the moon is only about one-sixth that of Earth. However, the rock contains the same amount of matter and thus has the same mass whether it is located on Earth or on the moon.

weight A measurement of the gravitational force acting on an object.

alsen/431 images/Pixabay

FIGURE 1.1 Objects on the moon would weigh about one-sixth of their weight on Earth.

Despite the difference in meaning between mass and weight, the determination of mass is commonly called "weighing." We will follow that practice in this book, but we will use the correct term *mass* when referring to an amount of matter.

1.2 Properties and Changes

Learning Objective 2 Explain the difference between the terms *physical* and *chemical* as applied to the properties of matter and changes in matter.

When you looked at your surroundings earlier, you didn't have much trouble identifying the various things you saw. For example, unless the decorator of your room had unusual tastes, you could easily tell the difference between a TV set and a potted plant by observing characteristics such as shape, color, and size. Our ability to identify objects or materials and discriminate between them depends on such characteristics. Scientists prefer to use the term *property* instead of *characteristic*, and they classify properties into two categories, physical and chemical.

physical properties Properties of matter that can be observed or measured without trying to change the composition of the matter being studied.

chemical properties Properties that matter demonstrates when attempts are made to change it into new substances.

Physical properties are those that can be observed or measured without changing or trying to change the composition of the matter i.e.—no original substances are destroyed, and no new substances appear. For example, you can observe the color or measure the size of a sheet of paper without attempting to change the paper into anything else. Color and size are physical properties (see **Figure 1.2**) of the paper. **Chemical properties** are the properties that matter demonstrates when attempts are made to change it into other kinds of matter. For example, a sheet of paper can be burned; in this process, the paper is changed into a new substance. On the other hand, attempts to burn a piece of glass under similar conditions fail. The ability of paper to burn is a chemical property, as is the inability of glass to burn.

physical changes Changes matter undergoes without changing composition.

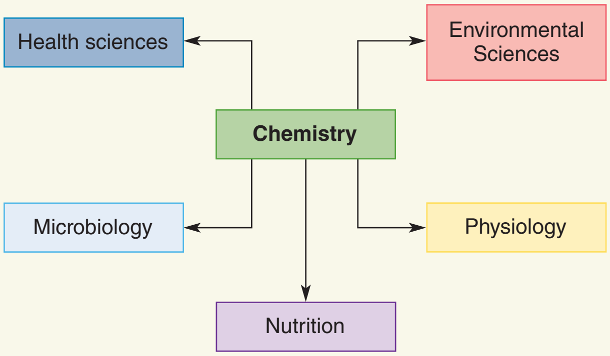
chemical changes Changes matter undergoes that involve changes in composition.

You can easily change the size of a sheet of paper by cutting off a piece. The paper sheet is not converted into any new substance by this change, but it is simply made smaller. **Physical changes** can be carried out without changing the composition of a substance. However, there is no way you can burn a sheet of paper without changing it into new substances. Thus, the change that occurs when paper burns is called a **chemical change. Figure 1.3** shows an example of a chemical change, the burning of magnesium metal. The bright light produced

ENVIRONMENTAL CONNECTIONS 1.1

A Central Science

Chemistry is often referred to as the "central science" because it serves as a necessary foundation for many other scientific disciplines. Regardless of which scientific field you are interested in, every single substance you will discuss or work with is made up of chemicals. Also, many processes important to those fields will be based on an understanding of chemistry.



Chemistry is the foundation for many other scientific disciplines.

We also consider chemistry a central science because of its crucial role in responding to the needs of society. We use chemistry to discover new processes, develop new sources of energy, produce new products and materials, provide more food, and ensure better health.

As you read this text, you will encounter chapter-opening applications of chemistry in the health-care professions. Within the chapters, Environmental Connections and Health Connections boxes focus on specific topics that play essential roles in meeting the needs of society.

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Chemicals are present in everything we can touch, smell, or see. Chemistry is all around us.

FIGURE 1.2 Some physical properties of matter.

by this chemical change led to the use of magnesium in the flash powder used in early photography. Magnesium is still used in fireworks to produce a brilliant white light.

Example 1.1 Classifying Changes as Physical or Chemical

Classify each of the following changes as physical or chemical: (a) a match is burned; (b) iron is melted; (c) limestone is crushed; (d) limestone is heated, producing lime and carbon dioxide; (e) an antacid seltzer tablet is dissolved in water; and (f) a rubber band is stretched.

Solution

Changes b, c, and f are physical changes because no composition changes occurred and no new substances were formed.

The others are chemical changes because new substances were formed as illustrated. A match is burned—combustion gases are given off, and matchstick wood is converted to ashes. Limestone is heated—lime and carbon dioxide are the new substances. A seltzer tablet is dissolved in water—the fizzing that results is evidence that at least one new material (a gas) is produced.

✔ **LEARNING CHECK 1.1** Classify each of the following changes as physical or chemical, and, in the cases of chemical change, describe one observation or test that indicates new substances have been formed: (a) milk sours, (b) a wet handkerchief dries, (c) fruit ripens, (d) a stick of dynamite explodes, (e) air is compressed into a steel container, and (f) water boils.

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1. A strip of magnesium metal.

2. After being ignited with a flame, the magnesium burns with a blinding white light.

3. The white ash of magnesium oxide from the burning of several magnesium strips.

FIGURE 1.3 A chemical change occurs when magnesium metal burns.

Among the most common physical changes are changes in state, such as the melting of solids to form liquids, the sublimation of solids to form gases, or the evaporation of liquids to form gases. These changes take place when heat is added to or removed from matter, as represented in **Figure 1.4**. We will discuss changes in state in more detail in Chapter 6.

FIGURE 1.4 Examples of physical change.

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a Solid iodine becomes gaseous iodine when heated.

b Liquid benzene becomes solid benzene when cooled.

1.3 A Model of Matter

Learning Objective 3 Describe matter in terms of the accepted scientific model.

Model building is a common activity of scientists, but the results in many cases would not look appropriate on a fireplace mantle. **Scientific models** are explanations for observed behavior. Some, such as the well-known representation of the solar system, can easily be depicted in a physical way. Others are so abstract that they can be represented only by mathematical equations.

scientific models Explanations for observed behavior in nature.

Our present understanding of the nature of matter is a model that has been developed and refined over many years. Based on careful observations and measurements of the properties of matter, the model is still being modified as more is learned. In this book, we will concern ourselves with only some very basic concepts of this model, but even then these basic ideas will provide a powerful tool for understanding the behavior of matter.

The study of the behavior of gases—such as air, oxygen, and carbon dioxide—by some of the earliest scientists led to a number of important ideas about matter. The volume of a gas kept at a constant temperature was found to change with pressure. An increase in pressure caused the gas volume to decrease, whereas a decrease in pressure permitted the gas volume to increase. It was also discovered that the volume of a gas maintained at constant pressure increased as the gas temperature was increased. Gases were also found to have mass and ability to mix rapidly with one another when brought together.


Alexandra Lande/Shutterstock.com

FIGURE 1.5 A hang glider soars far above the ground. How does this feat confirm that air is matter?

A simple model for matter was developed that explained these gaseous properties, as well as many properties of solids and liquids. Some details of the model are discussed in Chapter 6, but one conclusion is important to us now I.e., all matter is made up of particles that are too small to see (see **Figure 1.5**). The early framers of this model called the small particles *molecules*. It is now known that molecules are the constituent particles of many, but not in all substances. In this chapter, we will limit our discussion to substances made up of molecules. Substances that are not made of molecules will be discussed in **Sections 4.3** and **4.11**.

The results of some simple experiments will help us formally define the term *molecule*. Suppose you have a container filled with oxygen gas and you perform a number of experiments with it. You find that a glowing splinter of wood bursts into flames when placed in the gas. A piece of moist iron rusts much faster in the oxygen than it does in air. A mouse or other animal can safely breathe the gas.

Now suppose you divide another sample of oxygen, the same size as the first into two smaller samples. The results of similar experiments done with these samples would be the same as before. Continued subdivision of an oxygen sample into smaller and smaller samples does not change the ability of the oxygen in the samples to behave just like the oxygen in the original sample. We conclude that the physical division of a sample of oxygen gas into smaller and smaller samples does not change the oxygen into anything else—it is still oxygen. Is there a limit to such divisions? What is the smallest sample of oxygen that will behave like the larger sample? We hope you have concluded that the smallest sample must be a single molecule. Although its very small size would make a one-molecule sample difficult to handle, it would nevertheless behave just as a larger sample would—it could be stored in a container, it would make wood burn rapidly, it would rust iron, and it could be breathed safely by a mouse.

We are now ready to formally define the term *molecule*. A **molecule** is the smallest particle of a pure substance that has the properties of that substance and is capable of a stable independent existence. Alternatively, a molecule is defined as the limit of physical subdivision for a pure substance.

molecule The smallest particle of a pure substance that has the properties of that substance and is capable of a stable independent existence. Alternatively, a molecule is the limit of physical subdivision for a pure substance.

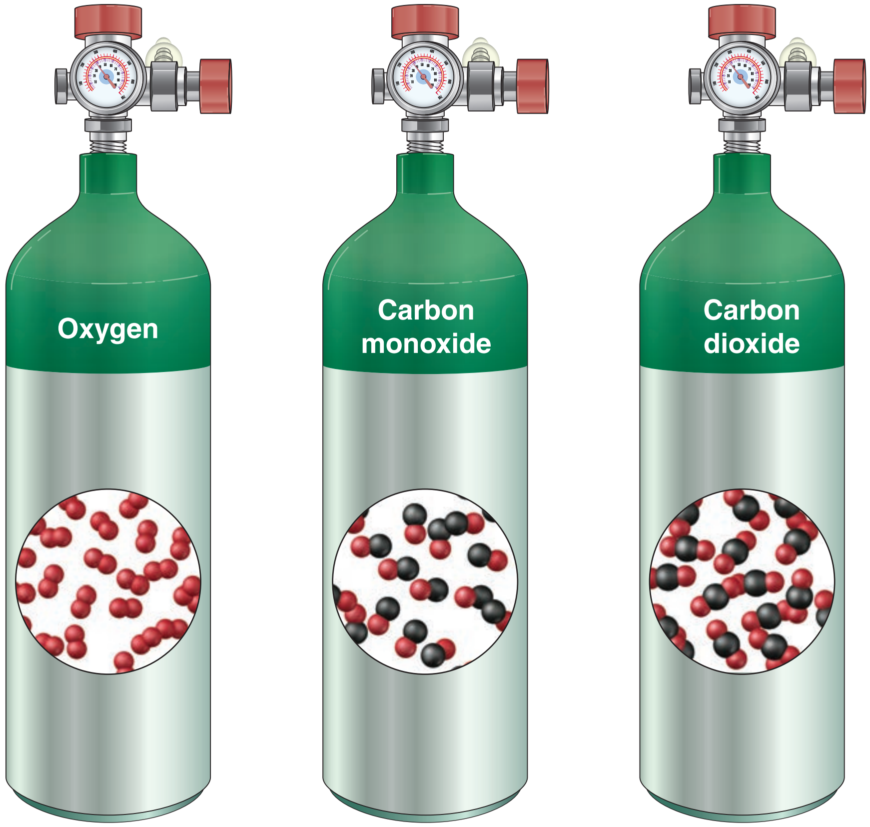
In less formal terms, these definitions indicate that a sample of pure substance—such as oxygen, carbon monoxide, or carbon dioxide—can be physically separated into smaller and smaller samples only until there is a single molecule. Any further separation cannot be done physically, but if it were done (chemically), the resulting sample would no longer have the same properties as the larger sample.

The idea that it might be possible to chemically separate a molecule into smaller particles grew out of continued study and experimentation by early scientists. In modern terminology, the smaller particles that make up molecules are called atoms. John Dalton (1766–1844) is generally credited with developing the first atomic theory containing ideas that are still used today. The main points of his theory, which he proposed in 1808, can be summarized in the following five statements:

1. All matter is made up of tiny particles called atoms.
2. Substances called elements are made up of atoms that are all identical.
3. Substances called compounds are combinations of atoms of two or more elements.
4. Every molecule of a specific compound always contains the same number of atoms of each kind of element found in the compound.
5. In chemical reactions, atoms are rearranged, separated, or combined, but are never created nor destroyed.

Early scientists used graphic symbols such as circles and squares to represent the few different atoms that were known at the time. Instead of different shapes, we will use representations such as those in **Figure 1.6** for oxygen, carbon monoxide, and carbon dioxide molecules.

FIGURE 1.6 Symbolic representations of molecules.



diatomic molecules Molecules that contain two atoms.

homoatomic molecules Molecules that contain only one kind of atom.

heteroatomic molecules Molecules that contain two or more kinds of atoms.

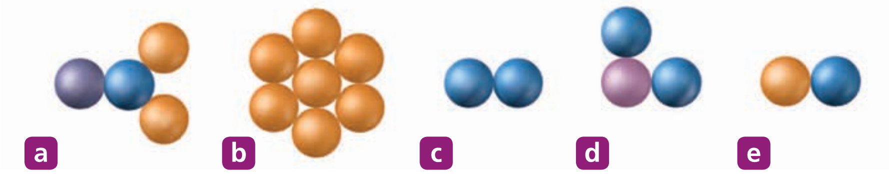
triatomic molecules Molecules that contain three atoms.

polyatomic molecules Molecules that contain three or more atoms.

The three pure substances just mentioned illustrate three different types of molecules found in matter. Oxygen molecules consist of two oxygen atoms, and are called **diatomic molecules** to indicate that fact. Molecules, such as oxygen, that contain only one kind of atom are also called **homoatomic molecules** to indicate that the atoms are all of the same kind. Carbon monoxide molecules also contain two atoms and therefore are diatomic molecules. However, in this case the atoms are not identical, a fact indicated by the term **heteroatomic molecule**. Carbon dioxide molecules consist of three atoms that are not all identical, so carbon dioxide molecules are described by the terms **triatomic** and heteroatomic. The words *diatomic* and *triatomic* are commonly used to indicate two and three atom molecules respectively, but the word **polyatomic** is usually used to describe molecules that contain more than three atoms.

Example 1.2 Classifying Molecules

Use the terms *diatomic, triatomic, polyatomic, homoatomic,* or *heteroatomic* to classify the following molecules correctly:



Solution

a. Polyatomic and heteroatomic (contains more than three atoms, and the atoms are not all identical)
b. Polyatomic and homoatomic (more than three atoms, and the atoms are all identical)
c. Diatomic and homoatomic (two atoms, and the atoms are identical)
d. Triatomic and heteroatomic (three atoms, and the atoms are not identical)
e. Diatomic and heteroatomic (two atoms, and the atoms are not identical)

✔ **LEARNING CHECK 1.2** Use the terms *diatomic, triatomic, polyatomic, homoatomic,* or *heteroatomic* to classify the following molecules correctly:

a. Water molecules have been found to contain two hydrogen atoms and one oxygen atom.
b. Molecules of ozone contain three oxygen atoms.
c. Natural gas is made up primarily of methane molecules, which contain one atom of carbon and four atoms of hydrogen.

The subdivision of molecules into smaller particles is a chemical change. How far can such subdivisions of molecules go? You are probably a step ahead of us and have guessed that the answer is atoms. In fact, this provides us with a definition of atoms. An **atom** is the limit of chemical subdivision. In less formal terms, atoms are the smallest particles of matter that can be produced as a result of chemical changes. However, all chemical changes do not necessarily break molecules into atoms. In some cases, chemical changes might just divide a large molecule into two or more smaller molecules. Also, as we will see later, some chemical changes form larger molecules from smaller ones. The important point is that only chemical changes will produce a division of molecules, and the smallest particles of matter that can possibly be produced by such a division are called atoms.

atom The limit of chemical subdivision for matter.

David Orcea/Shutterstock.com

FIGURE 1.7 A pure substance such as salt has a constant composition, 100%.

1.4 Classifying Matter

Learning Objective 4 On the basis of observation or information given to you, classify matter into the correct category of each of the following pairs: heterogeneous or homogeneous, solution or pure substance, and element or compound.

Unknown substances are often analyzed to determine their compositions. An analyst, upon receiving a sample to analyze, will always ask an important question: Is the sample a pure substance or a mixture? Any sample of matter must be one or the other. Pure water and sugar are both pure substances, but you can create a mixture by stirring together some sugar and pure water.

What is the difference between a pure substance and a mixture? Two differences are that a **pure substance** has a constant composition (see **Figure 1.7**) and a fixed set of physical and chemical properties. Pure water, for example, always contains the same proportions of hydrogen and oxygen and freezes at a specific temperature. A **mixture** of sugar and water, however, can vary in composition, and the properties will be different for

pure substance Matter that has a constant composition and fixed properties.

mixture A physical blend of matter that can theoretically be physically separated into two or more components.

different compositions. For example, a glass of sugar water could contain a few crystals of sugar or several spoonfuls. Properties such as the sweetness and freezing point would vary depending on the amount of sugar present in the mixture.

Another difference is that a pure substance cannot be physically separated into simpler substances, whereas a mixture can theoretically be separated into its constituent components. For example, if we heat a sugar-and-water mixture, the water evaporates, and the sugar remains there in. We say that mixtures can theoretically be separated, but some separations are very difficult to achieve. **Figure 1.8** summarizes these ideas.

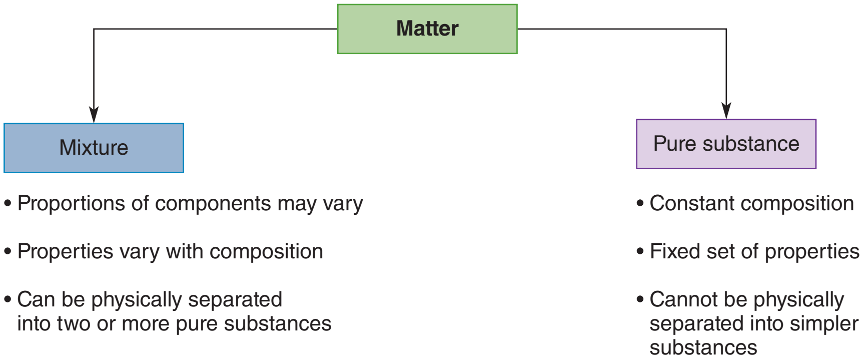


FIGURE 1.8 Mixtures and pure substances.

homogeneous matter Matter that has the same properties throughout the sample.

solutions Homogeneous mixtures of two or more pure substances.

heterogeneous matter Matter with properties that are not the same throughout the sample.

Pure substances, and mixtures such as sugar water, are examples of **homogeneous matter**—matter that has a uniform appearance and the same properties throughout. Homogeneous mixtures such as sugar water are called **solutions** (see **Figure 1.9**). Mixtures in which the properties and appearance are not uniform throughout the sample, are examples of **heterogeneous matter**. The mixture of rock salt and sand that is spread on snowy roads during the winter is an example.

Commonly, the word *solution* is used to describe homogeneous liquid mixtures such as sugar water, but solutions of gases and solids also exist. The air around us is a gaseous solution, containing primarily nitrogen and oxygen. The alloys of some metals are solid solutions. For example, small amounts of copper are often added to the gold used in making jewelry. The resulting solid solution is harder than gold and has greater resistance to wear.

Most matter is found in nature in the form of heterogeneous mixtures. The properties of such mixtures depend on the location from which their samples are taken. In some cases,

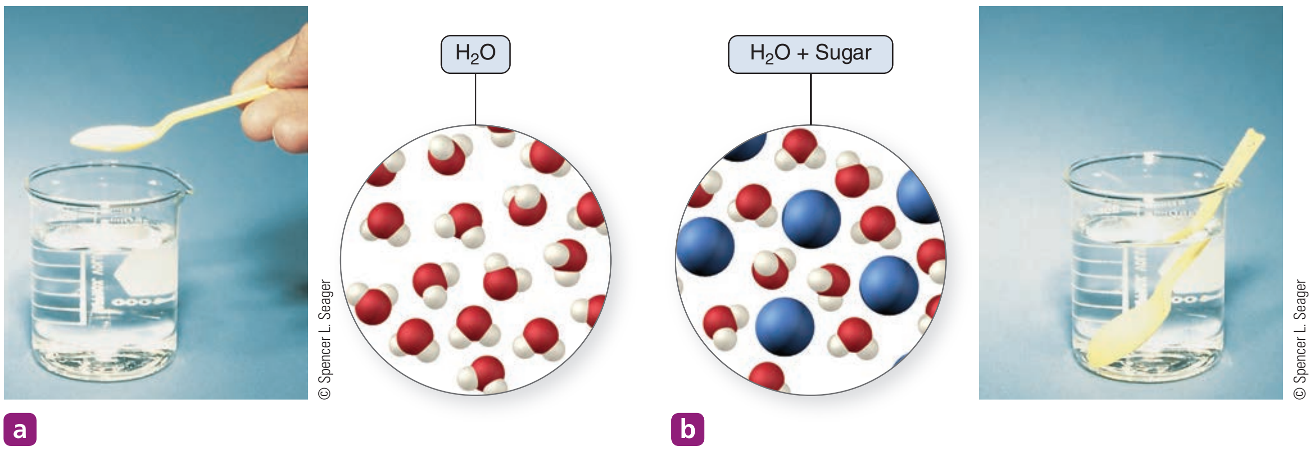


FIGURE 1.9 Sugar and water **(a)** form a solution when mixed **(b)**.

the heterogeneity is obvious. In a slice of tomato, for example, the parts representing the skin, juice, seeds, and pulp can be easily seen and identified, because they look different. Thus, at least one property (e.g., color or texture) is different for the different parts. However, a sample of clean sand collected from a seashore must be inspected very closely before slight differences in appearance can be seen for different sand particles. At this point, you might be thinking that even the solutions described earlier as homogeneous mixtures would appear to be heterogeneous if they were looked at closely enough. We could differentiate between sugar and water molecules if sugar solutions were observed under high magnification. We will generally limit ourselves to differences normally visible, when we classify matter as heterogeneous on the basis of appearance.

Earlier, we looked at three examples of pure substances—oxygen, carbon monoxide, and carbon dioxide—and found that the molecules of these substances are of different types. Oxygen molecules are diatomic and homoatomic, carbon monoxide molecules are diatomic and heteroatomic, and carbon dioxide molecules are triatomic and heteroatomic. Many pure substances have been found to consist of either homoatomic or heteroatomic molecules—a characteristic that permits them to be classified into one of two categories. Pure substances made up of homoatomic molecules are called **elements**, and those made up of heteroatomic molecules are called **compounds**. Thus, oxygen is an element, whereas carbon monoxide and carbon dioxide are compounds.

element A pure substance consisting of only one kind of atom in the form of homoatomic molecules or individual atoms.

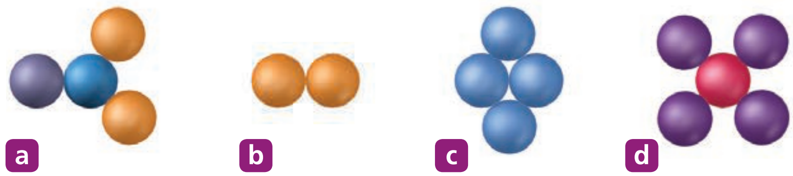
compound A pure substance consisting of two or more kinds of atoms in the form of heteroatomic molecules or individual atoms.

It is useful to note a fact here that is discussed in more detail in **Section 4.11**. The smallest particles of some elements and compounds are individual atoms rather than molecules. However, in elements of this type, the individual atoms are all of the same kind, whereas in compounds, two or more kinds of atoms are involved. Thus, the classification of a pure substance as an element or a compound is based on the fact that only one kind of atom is found in elements and two or more kinds are found in compounds. In both cases, the atoms may be present individually or in the form of homoatomic molecules (elements) or heteroatomic molecules (compounds). Some common household materials are pure substances (elements or compounds), such as aluminum foil, baking soda, table salt, etc. (see **Figure 1.10**).

Mike Slabaugh

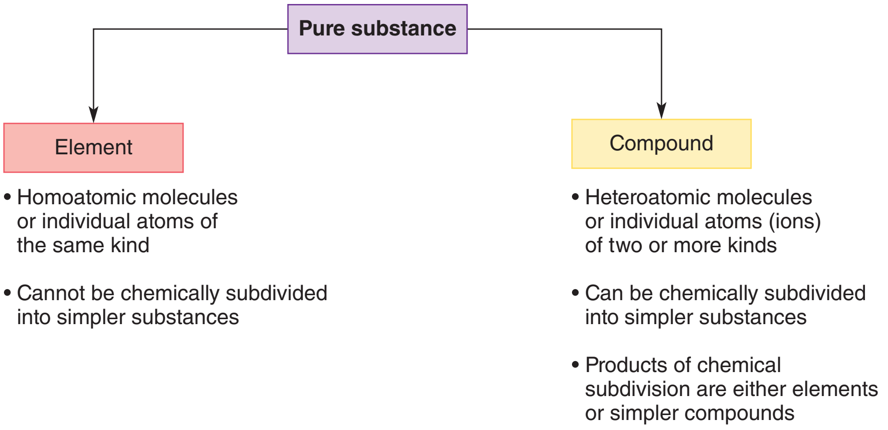
FIGURE 1.10 Sugar, baking soda, and aluminum foil are examples of pure substances found within the home.

✔ LEARNING CHECK 1.3 Classify the molecules represented below as those of an element or a compound:



The characteristics of the molecules of elements and compounds lead us to some conclusions about their chemical behavior. Elements cannot be chemically subdivided into simpler pure substances, but compounds can be. Because elements contain only one kind of atom and the atom is the limit of chemical subdivision. There is no chemical way to break an element into any simpler pure substance—the simplest pure substance is an element itself. On the other hand, because the molecules of compounds contain more than one kind of atom, breaking such molecules into simpler pure substances is possible. For example, a molecule of table sugar can be chemically changed into two simpler molecules (which are also sugars) or into atoms or molecules of the elements carbon, hydrogen, and oxygen. Thus, compounds can be chemically subdivided into simpler compounds or elements. **Figure 1.11** summarizes these ideas, and **Figure 1.12** illustrates a classification scheme for matter based on the ideas we have discussed.

FIGURE 1.11 Elements and compounds.



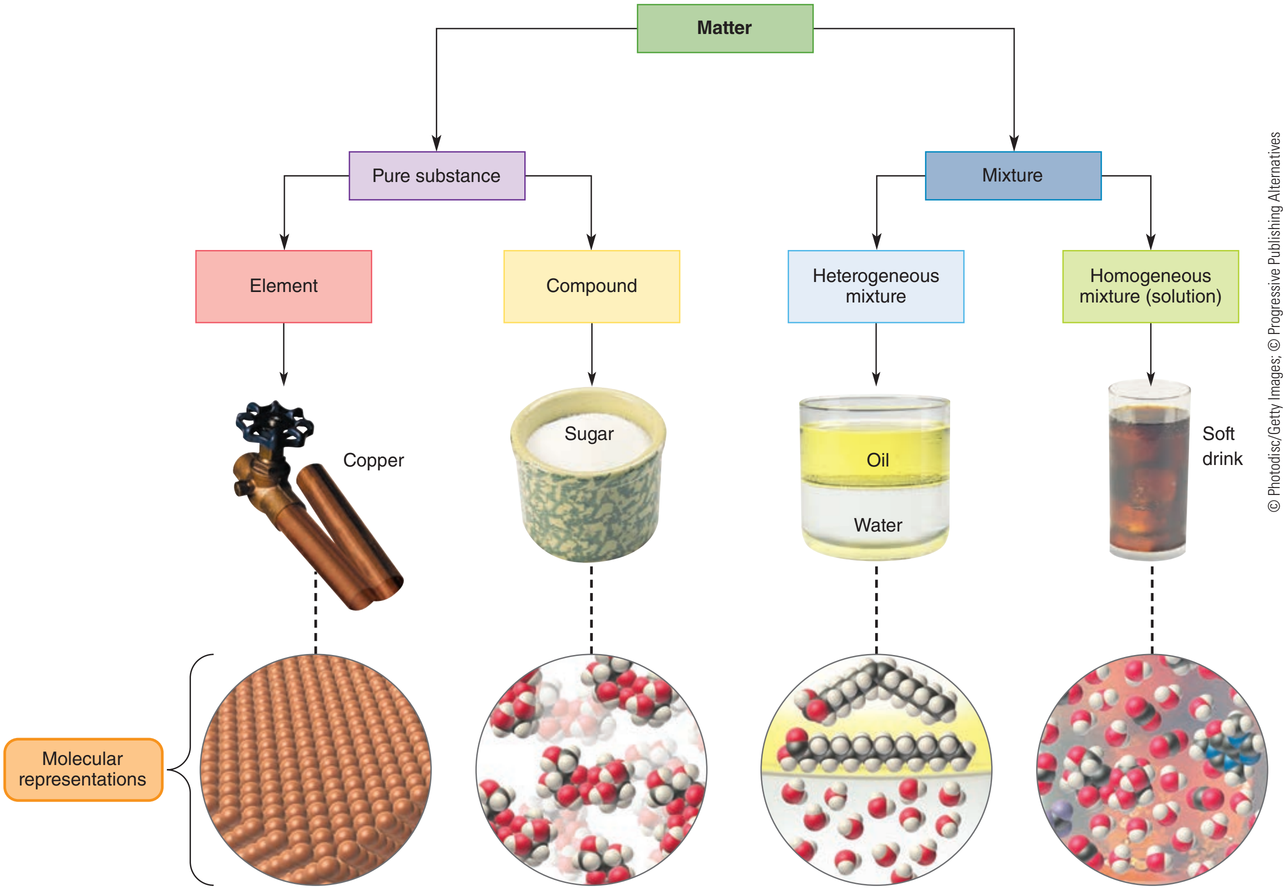


FIGURE 1.12 A classification scheme for matter.

Example 1.3 Classifying Substances

When sulfur, an element, is heated in air, it combines with oxygen to form sulfur dioxide. Classify whether sulphur dioxide is an element or a compound.

Solution

Because sulfur and oxygen are both elements and they combine to form sulfur dioxide, the molecules of sulfur dioxide must contain atoms of both sulfur and oxygen. Thus, sulfur dioxide is a compound because its molecules are heteroatomic.

✔ **LEARNING CHECK 1.4** Suppose an element and a compound combine to form only one product. Classify whether the product is an element or a compound.

1.5 Measurement Units

Learning Objective 5 Recognize the use of measurement units in everyday activities.

Matter can be classified and some physical or chemical properties can be observed without making any measurement. However, the use of quantitative measurements and calculations greatly expands our ability to understand the chemical nature of the world around us. A measurement consists of two parts, a number and an identifying unit. A number expressed without a unit is generally useless, especially in scientific work. We constantly make and express measurements in our daily lives. We measure the gallons of gasoline put into our cars, the time it takes to drive a certain distance, and the temperature on a hot or cold day. In some of our daily measurements, the units might be implied or understood. For example, if someone said the temperature outside was 39, you would probably assume this was 39 degrees Fahrenheit if you lived in the United States, but in most other parts of the world, it would be 39 degrees Celsius. Such confusion is avoided by expressing both the number and the unit of a measurement.

All measurements are based on units agreed on by those making and using the measurements. When a measurement is made in terms of an agreed-on unit, the result is expressed as some multiple of that unit. For example, when you purchase 10 pounds of potatoes, you are buying a quantity of potatoes equal to 10 times the standard quantity called 1 pound. Similarly, 3 feet of string is a length of string 3 times as long as the standard length that has been agreed on and called 1 foot.

The earliest units used for measurements were based on the dimensions of the human body. For example, the foot was the length of some important person's foot, and the biblical cubit (see **Figure 1.13**) was the length along the forearm from the elbow to the tip of the middle finger. One problem with such units is obvious; the size of the units changed when the person on whom they were based changed because of death, change in political power, and so on.

FIGURE 1.13 In ancient cultures, a cubit could vary from 18 to 21 inches, depending on the length of the arm.

As science became more quantitative, scientists found that the lack of standard units became more and more of a problem. A standard system of units was developed in France about the time of the French Revolution and was soon adopted by scientists throughout the world. This system, called the *metric system*, has since been adopted and is used by almost all nations of the world. The United States adopted the system but has not yet put it into widespread use.

In an attempt to further standardize scientific measurements, an international agreement in 1960 established certain basic metric units, and units derived from them, as preferred units to be used in scientific measurements. Measurement units in this system are known as SI units after the French Système International d'Unités. SI units have not yet been totally put into widespread use. Many scientists continue to express certain quantities, such as volume, in non-SI units. The metric system in this book is generally based on accepted SI units but also includes a few of the commonly used non-SI units.

1.6 The Metric System

Learning Objective 6 Recognize units of the metric system, and convert the measurements done, by using the metric system into related units.

The metric system has a number of advantages compared to other measurement systems. One of the most useful among all the advantages is that the metric system is a decimal system in which larger and smaller units of a quantity are related by factors of 10.

FIGURE 1.14 Some things that are about 1 meter in length.

basic unit of measurement A specific unit from which other units for the same quantity are obtained by multiplication or division.

derived unit of measurement A unit obtained by multiplication or division of one or more basic units.

See **Table 1.1** for a comparison between the metric and English units of length—a meter is slightly longer than a yard (see **Figure 1.14**). Notice in **Table 1.1** that the units of length in the metric system are related by multiplying a specific number by 10—remember, $100 = 10 \times 10$ and $1000 = 10 \times 10 \times 10$. The relationships between the units of the English system show no such pattern.

TABLE 1.1 Metric and English Units of Length

	Base Unit	Larger Unit	Smaller Unit
Metric	1 meter	1 kilometer = 1000 meters	10 decimeters = 1 meter 100 centimeters = 1 meter 1000 millimeters = 1 meter
English	1 yard	1 mile = 1760 yards	3 feet = 1 yard 36 inches = 1 yard

The relationships between units of the metric system that are larger or smaller than a **basic** (defined) **unit** are indicated by prefixes attached to the name of the basic unit. Thus, 1 kilometer (km) is a unit of length that is 1000 times longer than the basic unit of 1 meter (m), and a millimeter (mm) is only $\frac{1}{1000}$ the length of 1 m. Some commonly used prefixes are given in **Table 1.2**.

TABLE 1.2 Common Prefixes of the Metric System

Prefix[a]	Abbreviation	Relationship to Basic Unit	Exponential Relationship to Basic Unit[b]
mega-	M	1,000,000 × basic unit	10^{6} × basic unit
kilo-	k	1000 × basic unit	10^{3} × basic unit
deci-	d	1/10 × basic unit	10^{-1} × basic unit
centi-	c	1/100 × basic unit	10^{-2} × basic unit
milli-	m	1/1000 × basic unit	10^{-3} × basic unit
micro-	μ	1/1,000,000 × basic unit	10^{-6} × basic unit
nano-	n	1/1,000,000,000 × basic unit	10^{-9} × basic unit
pico-	p	1/1,000,000,000,000 × basic unit	10^{-12} × basic unit

[a]The prefixes in boldface (heavy) type are the most common ones. [b]The use of exponents to express large and small numbers is discussed in **Section 1.7**.

Area and volume are examples of **derived units** of measurement. They are obtained or derived from the basic unit of length as given below:

$$\text{area} = (\text{length})(\text{length}) = (\text{length})^2$$

$$\text{volume} = (\text{length})(\text{length})(\text{length}) = (\text{length})^3$$

The unit used to express an area or volume depends on the unit of length used.

Example 1.4 Calculating Areas

Calculate the area of a rectangle that has sides of 1.5 m and 2.0 m. Express the answer in units of square meters and square centimeters.

Solution

When working examples and exercises with calculations, it is important to use a stepwise approach. Learning the strategy here will help you in the future to deal with more challenging questions.

STEP 1 Examine the question for what is given and what is unknown.

Given: sides are 1.5 m and 2.0 m   Unknown: area = ?

STEP 2 Identify a formula that links the given and unknown terms.

$$\text{area} = (\text{length})(\text{length})$$

STEP 3 Enter the data into the formula.

$$\text{area} = (1.5\text{ m})(2.0\text{ m}) = 3.0\text{ m}^2$$

In **Table 1.1**, it is given that 1 m = 100 cm. In terms of centimeters, the lengths are 150 cm and 200 cm.

$$\text{area} = (150\text{ cm})(200\text{ cm}) = 30{,}000\text{ cm}^2$$

STEP 4 Check to see whether your answer makes sense.
Yes. The values seem to be in a reasonable range.

✔ **LEARNING CHECK 1.5** The area of a circle is given by the formula $A = \pi r^2$, where r is the radius and $\pi = 3.14$. Calculate the area of a circle that has a radius of 3.5 cm.

The unit used to express volume also depends on the unit of length used in the calculation. Thus, a volume could have such units as cubic meters (m^3), cubic decimeters (dm^3), or cubic centimeters (cm^3). The abbreviation cc is also used to represent cubic centimeters, especially in medical work. The liter (L), a non-SI unit of volume, has been used as a basic unit of volume by chemists for many years (see **Figure 1.15**). For all practical purposes, 1 L and 1 dm^3 are equal volumes. This also means that 1 milliliter (mL) is equal to 1 cm^3. Most laboratory glassware is calibrated in liters or milliliters.

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FIGURE 1.15 A liter is slightly larger than a quart.

Example 1.5 Expressing Volume in Metric Units

Laboratory tests for substances present in blood are often reported in mass per deciliter. From **Table 1.2**, we obtain the numerical relationship that deciliter (dL) = 0.1 liter (L). Express a volume of 1.36 L in terms of dL.

Solution

Because 1 dL = 0.1L, 1.36L can be converted to deciliters as

$$1.36\,\cancel{\text{L}} \times \frac{1\text{ dL}}{0.1\,\cancel{\text{L}}} = 13.6\text{ dL}$$

The units of the original quantity (1.36 L) were canceled, and the desired units (dL) were generated by the application of the factor-unit method (see **Section 1.9**).

The basic unit of mass in the metric system is 1 kilogram (kg), which is equal to about 2.2 pounds in the English system. A kilogram is too large to be conveniently used in some applications, so it is subdivided into smaller units. Two of these smaller units that are often used in chemistry are the gram (g) and milligram (mg) (see **Figure 1.16**). The prefixes *kilo-* (k) and *milli-* (m) indicate the following relationships between these units:

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

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

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

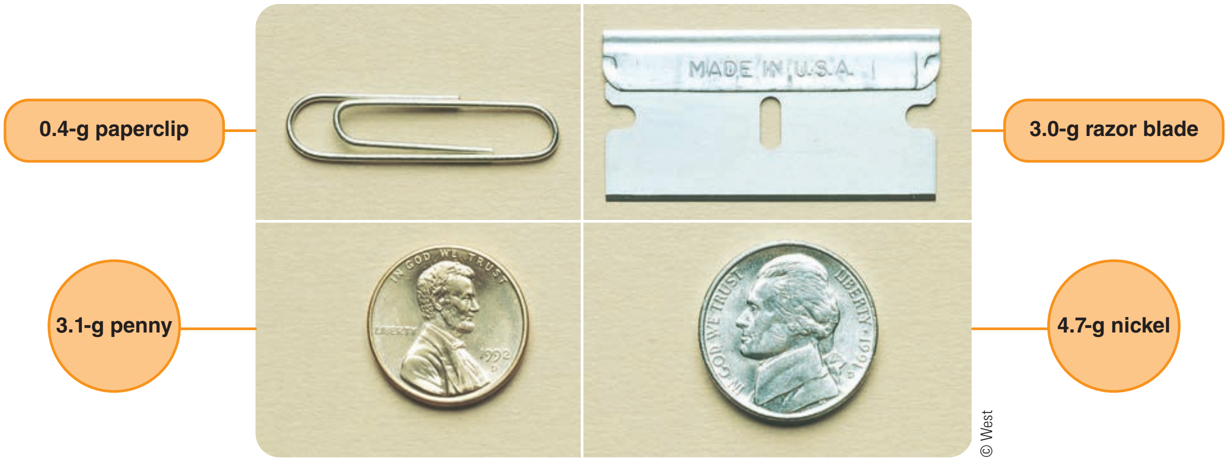


FIGURE 1.16 Metric masses of some common items as found in a 0.4-g paperclip, 3.0-g razor blade, 3.1-g penny, and 4.7-g nickel.

Example 1.6 Expressing Measurements in Metric Units

Saline solution for intravenous infusion is used to treat dehydration and has a number of other important uses in medicine. A partial bag of saline solution has a mass of 600 grams. Express this mass in kilograms and milligrams.

Solution

Since 1 kg = 1000 g, 600 g can be converted to kilograms as follows:

$$600\ \cancel{g} \times \frac{1\ \text{kg}}{1000\ \cancel{g}} = 0.600\ \text{kg}$$

Also, because 1 g = 1000 mg,

$$600\ \cancel{g} \times \frac{1000\ \text{mg}}{1\ \cancel{g}} = 600{,}000\ \text{mg}$$

Once again, the units of the original quantity (600 g) were canceled, and the desired units were generated by the application of the factor-unit method (see **Section 1.9**).

✔ **LEARNING CHECK 1.6** The javelin thrown by male competitors in track and field meets must have a minimum mass of 0.800 kg. A javelin is weighed and found to have a mass of 0.819 kg. Express the mass of the weighed javelin in grams.

Temperature is difficult to define but easy for most of us to measure—we just read a thermometer. However, thermometers can have temperature scales that represent different units. For example, a temperature of 293 would probably be considered quite high until it was pointed out that it is just room temperature as measured using the Kelvin temperature scale. Temperatures on this scale are given in kelvins, K. (Note that the abbreviation is K, not °K.)

The Celsius scale (formerly known as the centigrade scale) is the temperature scale used in most scientific work. On this scale, water freezes at 0°C and boils at 100°C under normal atmospheric pressure. A Celsius degree (division) is the same size as a kelvin of the Kelvin scale, but the two scales have different zero points. **Figure 1.17** compares the two scientific temperature scales and the familiar Fahrenheit scale. There are 100 Celsius degrees (divisions) between the freezing point (0°C) and the boiling point (100°C) of water. On the Fahrenheit scale, these same two temperatures are 180 degrees (divisions) apart

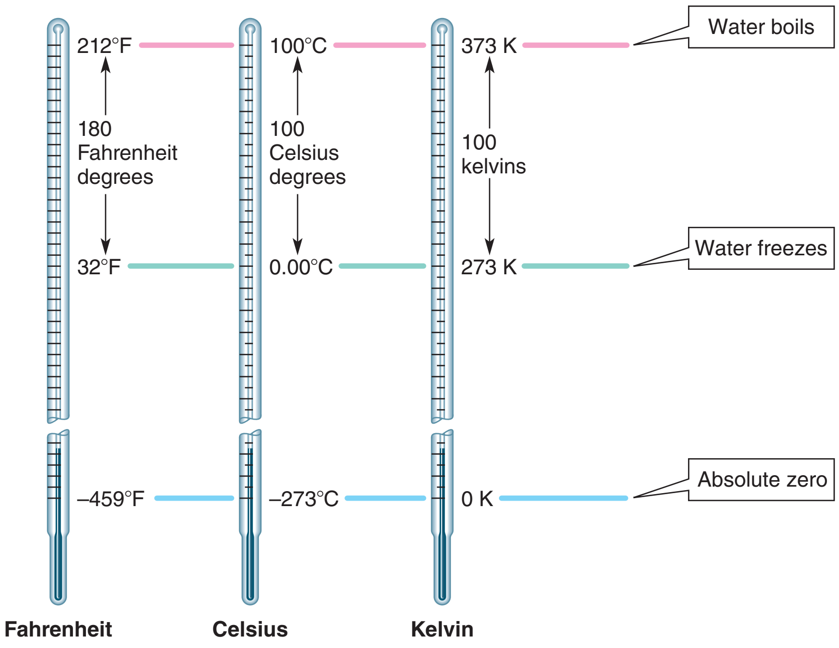


FIGURE 1.17 Fahrenheit, Celsius, and Kelvin temperature scales. The lowest temperature possible is absolute zero, 0 K.

(the freezing point is 32°F and the boiling point is 212°F). Readings on these two scales are related by the following equations:

$$°C = \frac{5}{9}(°F - 32) \tag{1.1}$$

$$°F = \frac{9}{5}(°C) + 32° \tag{1.2}$$

As mentioned, the difference between the Kelvin and Celsius scales is simply the zero point. Consequently, readings on the two scales are related as follows:

$$°C = K - 273 \tag{1.3}$$

$$K = °C + 273 \tag{1.4}$$

Notice that Equation 1.2 can be obtained by solving Equation 1.1 for Fahrenheit degrees, and Equation 1.4 can be obtained by solving Equation 1.3 for kelvins. Thus, you need to remember only Equations 1.1 and 1.3, rather than all four.

Example 1.7 Converting Fahrenheit and Celsius Temperatures

A mother is worried about her child, who has a temperature of 100.8°F. What reading would be shown if a Celsius thermometer were used?

Solution

STEP 1 Examine the question for what is given and what is unknown.

Given: 100.8°F

Unknown: °C = ?

STEP 2 Identify a formula that links the given and unknown terms.

$$\text{Key Equation 1.1: } °C = \frac{5}{9}(°F - 32)$$

STEP 3 Enter the data into the formula.

$$°C = \frac{5}{9}(100.8° - 32°) = \frac{5}{9}(68.8°) = 38.2°C$$

STEP 4 Check to see whether the answer makes sense.

Yes. °C is less than half of °F.

> ✔ **LEARNING CHECK 1.7** Body temperature is considered to be 37.0°C when measured orally. What reading would be shown if a Fahrenheit thermometer were used?

The last units discussed at this point are derived units of energy. Other units will be introduced later in the book as they are needed. The metric system unit of energy is a joule (J), pronounced "jewl." A joule is quite small, as shown by the fact that a 50-watt light-bulb uses 50 J of energy every second. A typical household in the United States uses several billion joules of electrical energy in a month.

The calorie (cal), a slightly larger unit of energy, is sometimes used by chemists. One calorie is the amount of heat energy required to increase the temperature of 1 g of water by 1°C. The calorie and joule are related as follows:

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

The nutritional calorie of the weight watcher is actually 1000 scientific calories, or 1 kcal. It is represented by writing *calorie* with a capital C (Calorie, abbreviated Cal). **Table 1.3** contains a list of the commonly used metric units, their relationship to basic units, and their relationship to English units.

TABLE 1.3 Commonly Used Metric Units

Quantity	Metric Unit	Relationship to Metric Basic Unit	Relationship to English Unit
Length	meter (m)	Basic unit	1 m = 1.094 yd
	centimeter (cm)	100 cm = 1 m	1 cm = 0.394 in.
	millimeter (mm)	1000 mm = 1 m	1 mm = 0.0394 in.
	kilometer (km)	1 km = 1000 m	1 km = 0.621 mi
Volume	cubic decimeter (dm^3)	Basic unit	1 dm^3 = 1.057 qt
	cubic centimeter (cm^3 or cc)	1000 cm^3 = 1 dm^3	1 cm^3 = 0.0338 fl oz
	liter (L)	1 L = 1 dm^3	1 L = 1.057 qt
	milliliter (mL)[a]	1000 mL = 1 dm^3	1 mL = 0.0338 fl oz
Mass	gram (g)	1000 g = 1 kg	1 g = 0.035 oz
	milligram (mg)	1,000,000 mg = 1 kg	1 mg = 0.015 grain
	kilogram (kg)	Basic unit	1 kg = 2.20 lb
Temperature	degree Celsius (°C)	1°C = 1 K	1°C = 1.80°F
	kelvin (K)	Basic unit	1 K = 1.80°F
Energy	calorie (cal)	1 cal = 4.184 J	1 cal = 0.00397 BTU[b]
	kilocalorie (kcal)	1 kcal = 4184 J	1 kcal = 3.97 BTU
	joule (J)	Basic unit	1 J = 0.000949 BTU
Time	second (s)	Basic unit	Same unit used

[a]1 mL = 1 cm^3. [b]A BTU (British thermal unit) is the amount of heat required to increase the temperature of 1 pound of water by 1°F.

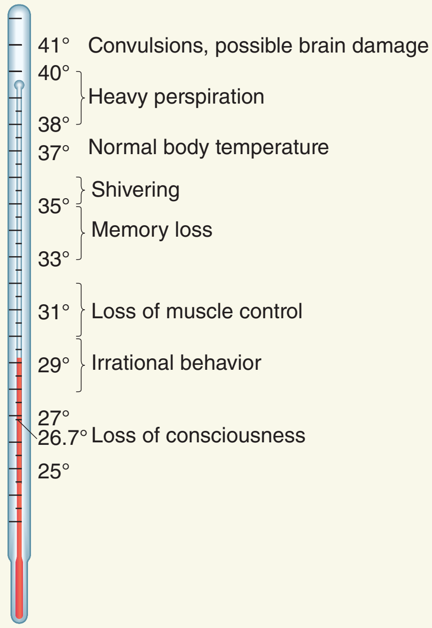
HEALTH CONNECTIONS 1.1

Effects of Temperature on Body Functions

The human body has the ability to remain at a relatively constant temperature even when the surrounding temperature increases or decreases. Because of this characteristic, we humans are classified as warm-blooded. In reality, our body temperature varies over a significant range depending on the time of day and the temperature of our surroundings. "Normal" body temperature is considered to be 37.0°C when measured orally. However, this "normal" value can fluctuate between a low of 36.1°C for an individual just waking up in the morning to a value as high as 37.2°C just before bedtime in the late evening.

In addition to this regular variation, our body temperature fluctuates in response to extremes in the temperature of our surroundings. In extremely hot environments, the capacity of our perspiration-based cooling mechanism can be overtaxed, and our body temperature will increase. Body temperatures more than 3.5°C above normal begin to interfere with bodily functions. Body temperatures above 41.1°C can cause convulsions and can result in permanent brain damage, especially in children.

Hypothermia occurs when the body's internal heat generation is not sufficient to balance the heat lost to very cold surroundings. As a result, the body temperature decreases, and at 28.5°C the afflicted person appears pale and might have an irregular heartbeat. Unconsciousness usually results if the body temperature gets lower than 26.7°C. At these low temperatures, respiration also slows down and becomes shallow, resulting in a decrease in the delivery of oxygen to the body tissues.



The effects of body temperature on body functions using the Celsius scale.

1.7 Large and Small Numbers

Learning Objective 7 Express numbers using scientific notation, and do calculations with numbers expressed in scientific notation.

Numbers are used in all measurements and calculations. Many numbers are readily understood and represented because of common experience with them. A price of 10 dollars, a height of 7 feet, a weight of 165 pounds, and a time of 40 seconds are examples of such numbers. But how do we handle numbers like the diameter of a hydrogen atom (about one hundred-millionth of a centimeter) or the distance light travels in 1 year—a light-year (about 6 trillion miles)? These numbers are so small and large, respectively, that they defy understanding in terms of relationships to familiar distances. Even if we can't totally relate to them, it is important in scientific work to be able to conveniently represent and work with such numbers.

Scientific notation provides a method for conveniently representing any number including those that are very large or very small. In **scientific notation**, numbers are represented as the product of a non-exponential term and an exponential term in the general form $M \times 10^n$. The non-exponential term M is a number between 1 and 10 (but not equal to 10) written with a decimal to the right of the first nonzero digit in the number. This position of the decimal is called the **standard position**. The exponential term is a 10 raised to a whole number exponent n that may be positive or negative. The value of n is the number

scientific notation A way of representing numbers consisting of a product between a nonexponential number and 10 raised to a whole-number exponent that may be positive or negative.

standard position for a decimal In scientific notation, the position to the right of the first nonzero digit in the nonexponential number.

of places the decimal must be moved from the standard position in M to be at the original position in the number when the number is written normally without using scientific notation. If n is positive, the original decimal position is to the right of the standard position [see **Figure 1.18(a)**]. If n is negative, the original decimal position is to the left of the standard position [see **Figure 1.18(b)**].

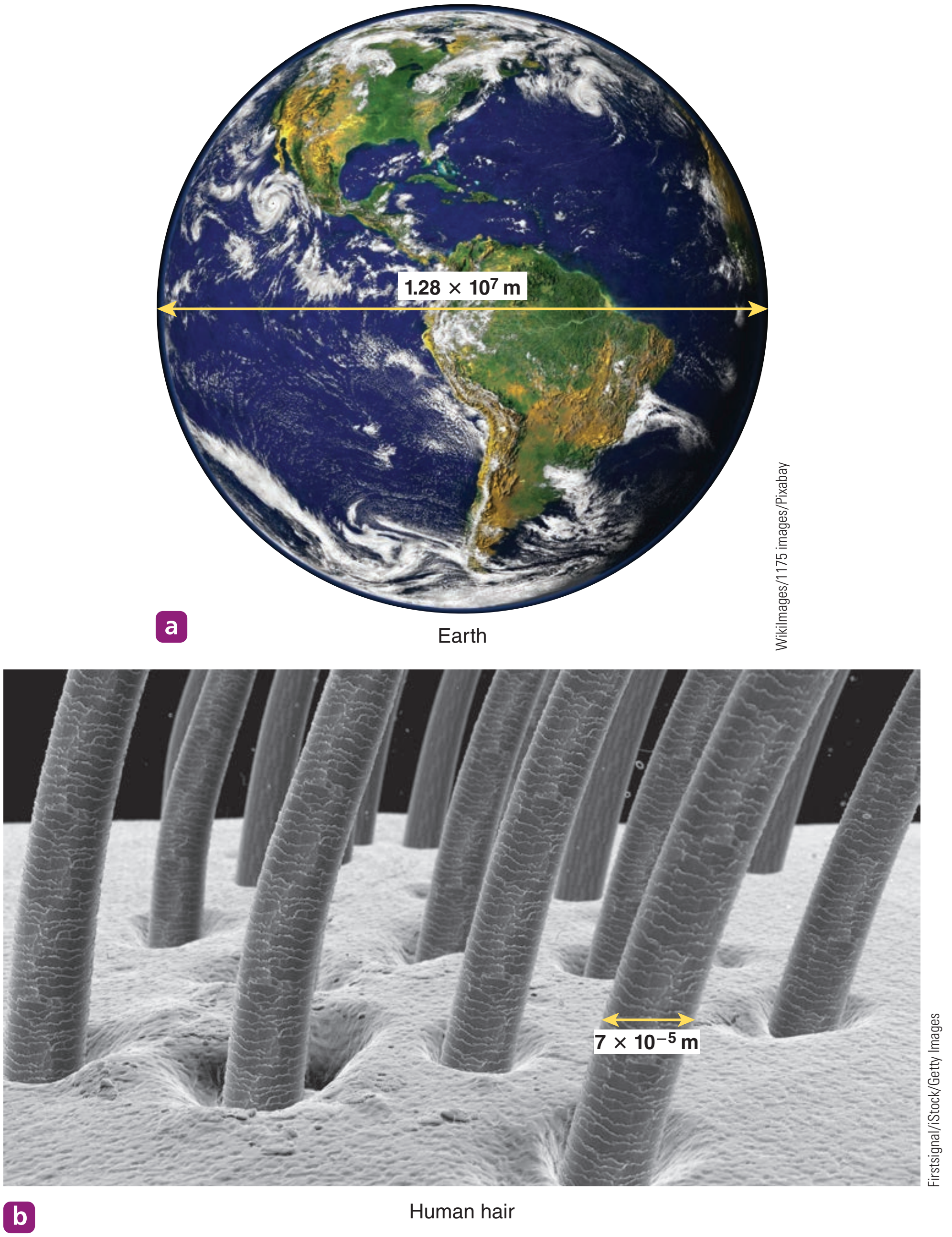


FIGURE 1.18 **(a)** The equatorial diameter of Earth is 1.28×10^7 meters. **(b)** The width of human hair varies depending on thickness. For average hair, the width is about 7×10^{-5} meters.

Example 1.8 Converting Scientific Notation to Non-scientific

Very small numbers such as the diameter of a bacterium and very large numbers such as the number of viral particles in a sample are expressed using scientific notation.

The following numbers are written using scientific notation. Write them without using scientific notation.

a. 3.72×10^5 b. 8.513×10^{-7}

Solution

a. The exponent 5 indicates that the new position of the decimal is located 5 places to the right of the standard position. Zeros are added to accommodate this change:

$$3.72 \times 10^5 = 372{,}000. = 372{,}000$$

standard position   new position

b. The exponent -7 indicates that the new position of the decimal is 7 places to the left of the standard position. Again, zeros are added as needed:

$$8.513 \times 10^{-7} = 0.0000008513$$

new position   standard position

Example 1.9 Converting Non-scientific Notation to Scientific

Write the following numbers using scientific notation:

a. 8725.6   b. 0.000729

Solution

a. The standard decimal position is between the 8 and 7: 8.7256. However, the original position of the decimal is 3 places to the right of the standard position. Therefore, the exponent must be $+3$:

$$8725.6 = 8.7256 \times 10^3$$

standard position   original position is to the right (+3)   exponent is positive for numbers greater than 1.0

b. The standard decimal position is between the 7 and 2: 7.29. However, the original position is 4 places to the left of standard. Therefore, the exponent must be -4:

$$0.000729 = 7.29 \times 10^{-4}$$

original position is to the left (−4)   standard position   exponent is negative for numbers less than 1.0

The zeros to the left of the 7 are dropped because they are not significant figures (see **Section 1.8**); they only locate the decimal in the nonscientific notation and are not needed in the scientific notation.

✔ **LEARNING CHECK 1.8** Some of the following numbers are written using scientific notation, and some are not. In each case, rewrite the numbers using the notation in which it is not written.

a. 5.88×10^2	c. 3.915×10^{-4}	e. 36.77
b. 0.000439	d. 9870	f. 0.102

Example 1.10 Recognizing Scientific Notation

Determine which of the following numbers are written correctly using scientific notation. For those that are not, rewrite them correctly.

a. 001.5×10^{-3} b. 28.0×10^{2} c. 0.35×10^{4}

Solution

a. Incorrect; the zeros to the left are not needed. The correct answer is 1.5×10^{-3}.
b. Incorrect; the decimal is not in the standard position. Move the decimal one position to the left and increase the exponent by 1 to give the correct answer of 2.80×10^{3}.
c. Incorrect; the decimal is not in the standard position. Move the decimal one position to the right and decrease the exponent by 1 to give the correct answer of 3.5×10^{3}.

✔ **LEARNING CHECK 1.9** Determine which of the following numbers are written correctly using scientific notation. For those that are not, rewrite them correctly.

a. 62.5×10^{4} b. 0.0098 c. 0.0041×10^{-3} d. 7.85×10^{2}

The multiplication and division of numbers written in scientific notation can be done quite simply by using some characteristics of exponentials. Consider the following multiplication:

$$(a \times 10^{y})(b \times 10^{z})$$

The multiplication is done in two steps. First, the nonexponential terms a and b are multiplied in the usual way. The exponential terms 10^{y} and 10^{z} are multiplied by adding the exponents y and z and using the resulting sum as a new exponent of 10. Thus, we can write

$$(a \times 10^{y})(b \times 10^{z}) = (a \times b)(10^{y+z})$$

Division is done similarly. The nonexponential terms are divided in the usual way, and the exponential terms are divided by subtracting the exponent of the bottom term from that of the top term. The final answer is then written as a product of the resulting nonexponential and exponential terms:

$$\frac{a \times 10^{y}}{b \times 10^{z}} = \left(\frac{a}{b}\right)(10^{y-z})$$

Multiplication and division calculations involving scientific notation are easily done using a calculator. **Table 1.4** gives the steps, the typical calculator procedures (buttons to press), and typical calculator readout or display for the division of 7.2×10^{-3} by 1.2×10^{4}.

TABLE 1.4 Using a Calculator for Scientific Notation Calculations

Step	Procedure		Calculator Display
1. Enter 7.2	Press buttons 7, ., 2	7 . 2	7.2
2. Enter 10^{-3}	Press button that activates exponential mode (EE, Exp, etc.)	EE , EXP	7.2^{00}
	Press 3	3	7.2^{03}
	Press change-sign button (±, etc.)	+/-	7.2^{-03}
3. Divide	Press divide button (÷)	÷	7.2^{-03}
4. Enter 1.2	Press buttons 1, ., 2	1 . 2	1.2
5. Enter 10^{4}	Press button that activates exponential mode (EE, Exp, etc.)	EE , EXP	1.2^{00}
	Press 4	4	1.2^{04}
6. Obtain answer	Press equals button (=)	=	$6.^{-07}$

Example 1.11 More Exercises with Scientific Notation

Do the following operations:

a. $(3.5 \times 10^4)(2.0 \times 10^2)$ c. $(4.6 \times 10^{-7})(5.0 \times 10^3)$

b. $\dfrac{3.8 \times 10^5}{1.9 \times 10^2}$ d. $\dfrac{1.2 \times 10^3}{3.0 \times 10^{-2}}$

Solution

a. $(3.5 \times 10^4)(2.0 \times 10^2) = (3.5 \times 2.0)(10^4 \times 10^2)$
$= (7.0)(10^{4+2}) = 7.0 \times 10^6$

b. $\dfrac{3.8 \times 10^5}{1.9 \times 10^2} = \dfrac{3.8}{1.9} \times \dfrac{10^5}{10^2} = (2.0)(10^{5-2}) = 2.0 \times 10^3$

c. $(4.6 \times 10^{-7})(5.0 \times 10^3) = (4.6 \times 5.0)(10^{-7} \times 10^3)$
$= (23)(10^{-7+3}) = 23 \times 10^{-4}$

To get the decimal into the standard position, move it one place to the left. This changes the exponent from -4 to -3, so the final result is 2.3×10^{-3}. (This number in decimal form, 0.0023, can be written correctly as either 23×10^{-4} or 2.3×10^{-3}, but scientific notation requires that the decimal point should be to the right of the first nonzero number.)

d. $\dfrac{1.2 \times 10^3}{3.0 \times 10^{-2}} = \dfrac{1.2}{3.0} \times \dfrac{10^3}{10^{-2}} = (0.40)(10^{3-(-2)}) = 0.40 \times 10^5$

Adjust the decimal to the standard position and get 4.0×10^4. If these examples were done using a calculator, the displayed answers would normally be given in correct scientific notation.

✔ **LEARNING CHECK 1.10** Perform the following operations, and express the result in correct scientific notation:

a. $(2.4 \times 10^3)(1.5 \times 10^4)$ c. $\dfrac{6.3 \times 10^5}{2.1 \times 10^3}$

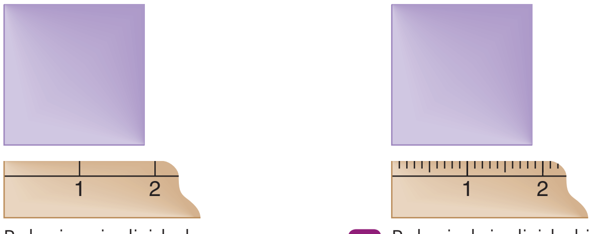
b. $(3.5 \times 10^2)(2.0 \times 10^{-3})$ d. $\dfrac{4.4 \times 10^{-2}}{8.8 \times 10^{-3}}$

The diameter of a hydrogen atom, mentioned earlier as one hundred-millionth of a centimeter, is written in scientific notation as 1.0×10^{-8} cm. Similarly, 1 light-year of 6 trillion miles is 6.0×10^{12} mi.

1.8 Significant Figures

Learning Objective 8 Express the results of measurements and calculations using the correct number of significant figures.

Every measurement contains an uncertainty that is characteristic of the device used to make the measurement. These uncertainties are represented by the numbers used to record the measurement. Consider the small square in **Figure 1.19**:



a Ruler in a is divided into centimeters.

b Ruler in b is divided into tenths of centimeters.

FIGURE 1.19 Measurements using two different rulers.

In (a), the length of one side of the square is measured with a ruler divided into centimeters. It is easy to see that the length is greater than 1 cm, but not quite 2 cm. The length is recorded by writing the number that is known with certainty to be correct (the 1) and writing an estimate for the uncertain number. The result is 1.9 cm, where the .9 is the estimate. In (b), the ruler is divided into tenths of centimeters. It is easy to see that the length is at least 1.8 cm, but not quite 1.9 cm. Once again the certain numbers (1.8) are written, and an estimate is made for the uncertain part. The result is 1.86 cm. When measurements are recorded this way, the numbers representing the certain measurement plus one number representing the estimate are called **significant figures**. Thus, the first measurement of 1.9 cm contains two significant figures, and the second measurement of 1.86 cm contains three significant figures.

significant figures The numbers in a measurement that represent the certainty of the measurement, plus one number representing an estimate.

The maximum number of significant figures possible in a measurement is determined by the design of the measuring device and cannot be changed by expressing the measurement in different units. The 1.8 cm length determined earlier can also be represented in terms of meters and millimeters as follows:

$$1.8 \text{ cm} = 0.018 \text{ m} = 18 \text{ mm}$$

In this form, it appears that the length expressed as 0.018 m contains four significant figures, but this is impossible; a measurement made with a device doesn't become more certain simply by changing the unit used to express the number. Thus, the zeros are not significant figures, their only function is to locate the correct position for the decimal. Zeros located to the left of nonzero numbers, such as the two zeros in 0.018 cm, are never considered to be significant. Thus, 12.5 mg, 0.0125 g, and 0.0000125 kg all represent the same measured mass, and all contain three significant figures.

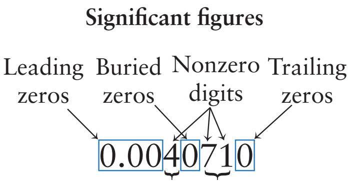


FIGURE 1.20 Five significant figures.

Zeros located between nonzero numbers are considered significant. Thus, 2055 μL, 2.055 mL, and 0.002055 L all represent the same volume measurement, and all contain four significant figures. Trailing zeros at the end of a number are significant if there is a decimal point present in the number. Otherwise, trailing zeros are not significant (see **Figure 1.20**).

The rule that covers trailing zeros is generally followed by scientists, but sometimes there is ambiguity. For example, suppose you read in a newspaper that the population of a city is 1,250,000 people. Should the four trailing zeros be considered significant? If they are, it means that the population is known with certainty to the nearest 10 people and that the measurement has an uncertainty of only plus or minus 1 person. A more reasonable conclusion is that the census is correct to the nearest 1000 people. This could be represented using scientific notation, 1.250×10^6. In this case, only four significant figures are used. In scientific notation, the correct number of significant figures is used in the nonexponential term, and the location of the decimal is determined by the exponent. In this book, large numbers will always be represented by scientific notation. However, you are likely to encounter nonsignificant trailing zeros in other reading materials. The rules for determining the significance of zeros are summarized as follows:

1. Zeros not preceded by nonzero numbers are not significant figures. These zeros are sometimes called *leading zeros.*
2. Zeros located between nonzero numbers are significant figures. These zeros are sometimes called *buried* or *confined zeros.*
3. Zeros located at the end of a number are significant figures if there is a decimal point present in the number. These zeros are sometimes called *trailing zeros.* If no decimal point is present, trailing zeros are not significant.

Example 1.12 Significant Figures and Scientific Notation

Determine the number of significant figures in each of the following measurements, and use scientific notation to express each measurement using the correct number of significant figures:

a. 24.6°C b. 0.036 g c. 15.0 mL d. 0.0020 m

Solution

a. All the numbers are significant: three significant figures, 2.46×10^1 °C.
b. The leading zeros are not significant: two significant figures, 3.6×10^{-2} g.
c. The trailing zero is significant: three significant figures, 1.50×10^1 mL.
d. The leading zeros are not significant, but the trailing zero is: two significant figures, 2.0×10^{-3} m.

✔ **LEARNING CHECK 1.11** Determine the number of significant figures in each of the following measurements:

a. 250 mg c. 0.0108 kg e. 0.001 mm
b. 18.05 mL d. 37°C f. 101.0 K

✔ **LEARNING CHECK 1.12** Use scientific notation to express each of the following measurements using the correct number of significant figures:

a. 101 m c. 0.00230 kg e. 21.65 mL
b. 1200 g d. 1296°C f. 0.015 km

Most measurements that are made do not stand as final answers. Instead, they are usually used to make calculations involving multiplication, division, addition, or subtraction. The answer obtained from such a calculation cannot have more certainty than the least certain measurement used in the calculation. It should be written to reflect an uncertainty equal to that of the most uncertain measurement. This is accomplished by the following rules:

1. The answer obtained by multiplication or division must contain the same number of significant figures as the quantity with the fewest significant figures used in the calculation.
2. The answer obtained by addition or subtraction must contain the same number of places to the right of the decimal as the quantity in the calculation with the fewest number of places to the right of the decimal.

To follow these rules, it is often necessary to reduce the number of significant figures by rounding answers. The following are rules for rounding:

1. If the first of the nonsignificant figures to be dropped from an answer is 5 or greater, all the nonsignificant figures are dropped, and the last significant figure is increased by 1.
2. If the first of the nonsignificant figures to be dropped from an answer is less than 5, all nonsignificant figures are dropped, and the last significant figure is left unchanged.

Remember, if you use a calculator, it will often express answers with too few or too many figures (see **Figure 1.21**). It will be up to you to determine the proper number of significant figures to use and to round the calculator answer correctly.

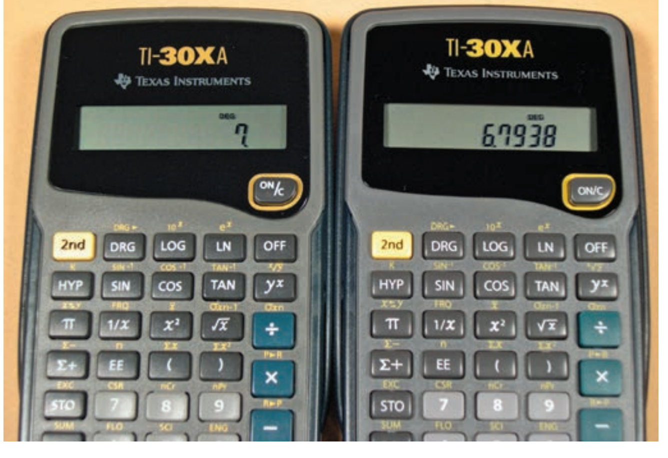


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FIGURE 1.21 Calculators usually display the sum of 4.362 and 2.638 as 7 (too few figures), and the product of 0.67 and 10.14 as 6.7938 (too many figures).

Example 1.13 Significant Figures in Multiplication and Division

On a typical day, nurses make numerous measurements and calculations. Reporting data in the most meaningful way is critical. Do the following calculations, and round the answers to the correct number of significant figures:

a. $(4.95)(12.10)$ b. $\dfrac{3.0954}{0.0085}$ c. $\dfrac{(9.15)(0.9100)}{3.117}$ d. $\dfrac{320}{4.00}$

Solution

All calculations are done with a calculator, and the calculator answer is written first. Appropriate rounding is done to get the final answer.

a. Calculator answer: 59.895
The number 4.95 has three significant figures, and 12.10 has four. Thus, the answer must have three significant figures:

$$\underbrace{59.8}_{\substack{\text{significant}\\ \text{figures}}}\underbrace{95}_{\substack{\text{nonsignificant}\\ \text{figures}}}$$

The first of the nonsignificant figures to be dropped is 9, so after both are dropped, the last significant figure is increased by 1. The final answer containing three significant figures is 59.9.

b. Calculator answer: 364.16471
The number 3.0954 has five significant figures, and 0.0085 has two. Thus, the answer must have only two:

$$\underbrace{36}_{\substack{\text{significant}\\ \text{figures}}}\underbrace{4.16471}_{\substack{\text{nonsignificant}\\ \text{figures}}}$$

The first of the nonsignificant figures to be dropped is 4, so the last significant figure remains unchanged after the nonsignificant figures are dropped. The correct answer then is 360, which contains two significant figures. The answer also can be written with the proper number of significant figures by using scientific notation, 3.6×10^2.

c. Calculator answer: 2.6713186
The number 9.15 has three significant figures, 0.9100 has four, and 3.117 has four. Thus, the answer must have only three:

$$\underbrace{2.67}_{\substack{\text{significant}\\ \text{figures}}}\underbrace{13186}_{\substack{\text{nonsignificant}\\ \text{figures}}}$$

Appropriate rounding gives 2.67 as the correct answer.

d. Calculator answer: 80
The number 320 has two significant figures and 4.00 contains three significant figures, so the answer should have two. The correct answer should be expressed as 80.0, which contains two significant figures.

✔ **LEARNING CHECK 1.13** Do the following calculations, and round the answers to the correct number of significant figures:

a. $(0.0019)(21.39)$ b. $\dfrac{8.321}{4.1}$ c. $\dfrac{(0.0911)(3.22)}{1.379}$

Example 1.14 Significant Figures in Addition and Subtraction

Do the following additions and subtractions, and write the answers with the correct number of significant figures:

a. $1.9 + 18.65$ b. $15.00 - 8.0$ c. $1500 + 10.9 + 0.005$ d. $5.1196 - 5.02$

Solution

In each case, the numbers are arranged vertically with the decimals in a vertical line. The answer is then rounded so it contains the same number of places to the right of the decimal as the smallest number of places in the quantities added or subtracted.

a.
$$\begin{array}{r} 1.9 \\ \underline{18.65} \\ 20.55 \end{array}$$

The answer must be expressed with one place to the right of the decimal to match the one place in 1.9.

Correct answer: 20.6 (Why was the final 5 increased to 6?)

b.
$$\begin{array}{r} 15.00 \\ \underline{-8.0} \\ 7.0 \end{array}$$

The answer must be expressed using one place to the right of the decimal to match the 8.0. A typical calculator answer would probably be 7, requiring that a zero be added to the right of the decimal to provide the correct number of significant figures.

c.
$$\begin{array}{r} 1500. \\ 10.9 \\ \underline{0.005} \\ 1510.905 \end{array}$$

The answer must be expressed with no places to the right of the decimal to match the 1500.

Correct answer: 1511 (Why was the final 0 of the answer increased to 1?)

d.
$$\begin{array}{r} 5.1196 \\ \underline{-5.02} \\ 0.0996 \end{array}$$

The answer must be expressed with two places to the right of the decimal to match the 5.02.

Correct answer: 0.10 (Which rounding rule was followed?)

Notice that the answer to (a) has more significant figures than the least significant number used in the calculation. The answer to (d), on the other hand, has fewer significant figures than either number used in the calculation. This happened because the rule for dealing with addition and subtraction focuses on the number of figures located to the right of the decimal and is not concerned with the figures to the left of the decimal. Thus, the number of significant figures in the answer is sometimes increased as a result of addition and decreased as a result of subtraction. You should be aware of this and not be confused when ever it happens.

✔ **LEARNING CHECK 1.14** Do the following additions and subtractions, and round the answers to the correct number of significant figures:

a. $8.01 + 3.2$
b. $3000 + 20.3 + 0.009$
c. $4.33 - 3.12$
d. $6.023 - 2.42$

Some numbers used in calculations are **exact numbers** that have no uncertainty associated with them and are considered to contain an unlimited number of significant trailing zeros. Such numbers are not used when the appropriate number of significant figures is determined for calculated answers. In other words, exact numbers do not limit the number of significant figures in calculated answers. One kind of exact number is a number used as part of a defined relationship between quantities. For example, 1 m contains exactly 100 cm:

$$1 \text{ m} = 100 \text{ cm}$$

Thus, the numbers 1 and 100 are exact. A second kind of exact number is a counting number obtained by counting individual objects. A dozen eggs contains exactly 12 eggs, not

exact numbers Numbers that have no uncertainty; numbers from defined relationships, counting numbers, and numbers that are part of simple fractions.

11.8 and not 12.3. The 12 is an exact counting number. A third kind of exact number is one that is part of a simple fraction such as $\frac{1}{4}$, $\frac{2}{3}$, or $\frac{5}{9}$ used in Equation 1.1 to convert Fahrenheit temperature readings into Celsius readings.

1.9 Using Units in Calculations

Learning Objective 9 Use the factor-unit method to solve numerical problems.

Some beginning chemistry students are concerned about not being able to solve numerical chemistry problems. They may say, "I can work the problems, I just can't set them up." What they are really saying is "I can do the arithmetic once the numbers are properly arranged, but I can't do the arranging."

This section presents a method for arranging numbers that will work for most of the numerical problems you will encounter in this course. This method has a number of names, including the factor-unit method, the factor-label method, and dimensional analysis. We will call it the factor-unit method. It is a systematic approach to solve numerical problems and consists of the following steps:

STEP 1 Write down the known or given quantity. Include both the numerical value and units of the quantity.

STEP 2 Leave some working space and set the known quantity equal to the units of the unknown quantity.

STEP 3 Multiply the known quantity by one or more factors, such that the units of the factor cancel the units of the known quantity and generate the units of the unknown quantity. These **factors** are fractions derived from numerical relationships between quantities. These relationships can be definitions or experimentally measured quantities. For example, the defined relationship 1 m = 100 cm provides the following two factors:

factors used in the factor-unit method Fractions obtained from numerical relationships between quantities.

$$\frac{1 \text{ m}}{100 \text{ cm}} \quad \text{and} \quad \frac{100 \text{ cm}}{1 \text{ m}}$$

STEP 4 After you get the desired units, do the necessary arithmetic to produce the final answer.

Example 1.15 Factor-Unit Calculations

As part of an annual physical exam, Brent's weight was recorded as 145 lb. What would Brent's weight be in kilograms? Use the factor-unit method and numerical relationships from **Table 1.3**.

Solution

The known quantity is 145 lb, and the unit of the unknown quantity is kilograms (kg).

STEP 1 145 lb

STEP 2 145 lb = ___ kg

STEP 3 $145 \cancel{\text{lb}} \times \dfrac{1 \text{ kg}}{2.20 \cancel{\text{lb}}} = \quad \text{kg}$

The factor $\dfrac{1 \text{ kg}}{2.20 \text{ lb}}$ came from the numerical relationship

1 kg = 2.20 lb found in **Table 1.3**.

STEP 4 $\dfrac{(145)(1) \text{ kg}}{2.20} = 65.9090 \text{ kg} = 65.9 \text{ kg}$

STEP 5 Check to see whether the answer makes sense.

Yes. The answer should be rounded to 65.9 kg, an answer that contains three significant figures, just as 145 lb does. The 1 kg in the factor is an exact number used as part of a defined relationship, so it doesn't influence the number of significant figures in the answer.

Example 1.16 More Factor-Unit Calculations

A clinical technician uses a micropipet to measure a 5.0 μL (5.0 microliter) sample of blood serum for analysis. Express the sample volume in liters (L).

Solution

The known quantity is 5.0 μL, and the unit of the unknown quantity is liters.

STEP 1 5.0 μL

STEP 2 $5.0\ \mu\text{L} \times \underline{\hspace{3cm}} = \text{L}$

STEP 3 $5.0\ \cancel{\mu\text{L}} \times \dfrac{1 \times 10^{-6}\ \text{L}}{1\ \cancel{\mu\text{L}}} = \text{L}$

The factor $\dfrac{1 \times 10^{-6}\ \text{L}}{1\ \mu\text{L}}$ came from the numerical relationship $1\ \mu\text{L} = 1 \times 10^{-6}\ \text{L}$ described in **Table 1.2**.

STEP 4 $\dfrac{(5.0)(1 \times 10^{-6}\ \text{L})}{1} = 5.0 \times 10^{-6}\ \text{L}$

The answer is expressed using the same number of significant figures as the 5.0 μL because 1 and 1×10^{-6} are exact numbers by definition.

STEP 5 Check to see whether your answer makes sense.

Yes. The unit is correct and the number seems reasonable.

✔ **LEARNING CHECK 1.15** Creatinine is a substance found in the blood. An analysis of a blood serum sample detected 1.1 mg of creatinine. Express this amount in grams by using the factor-unit method. Remember, the prefix milli- means $\frac{1}{1000}$, so $1\ \text{g} = 1000\ \text{mg}$.

Example 1.17 More Factor-Unit Calculations

One of the fastest-moving nerve impulses in the body travels at a speed of 405 feet per second (ft/s). What is the speed in miles per hour?

Solution

The known quantity is 405 ft/s, and the unit of the unknown quantity is miles per hour (mi/h).

STEP 1 $\dfrac{405\ \text{ft}}{\text{s}}$

STEP 2 $\dfrac{405\ \text{ft}}{\text{s}} \qquad = \dfrac{\text{mi}}{\text{h}}$

STEP 3 $\left(\dfrac{405\ \cancel{\text{ft}}}{\cancel{\text{s}}}\right)\left(\dfrac{1\ \text{mi}}{5280\ \cancel{\text{ft}}}\right)\left(\dfrac{60\ \cancel{\text{s}}}{1\ \cancel{\text{min}}}\right)\left(\dfrac{60\ \cancel{\text{min}}}{1\ \text{h}}\right) = \dfrac{\text{mi}}{\text{h}}$