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# Scientific Measurements

## This Chapter in Context

In the previous chapter, we saw that chemistry is an atomic and molecular science. The more we understand the nature of our atoms and molecules on the microscopic scale, the better we will be able to understand the properties of chemicals on the macroscopic scale; that is, the scale of things that we see and interact with. We obtain this understanding of the microscopic scale by making precise observations and numerical measurements. One place that measurements are made is in a bakery, to make different types of breads, different ratios of flour, water, and yeast are mixed. Then they are allowed to rise for different lengths of time, and finally, they are baked at different temperatures. All of these measurements together result in the different breads. In order to study chemistry, scientists systematically make and report measurements that allow them to explain the microscopic world.

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### CHAPTER OUTLINE



### LEARNING OBJECTIVES

**After reading, studying, and understanding this chapter, you should be able to:**

- explain the scientific method, and define law and theory.
- classify matter.
- identify which physical and chemical properties are important to measure.
- describe measurements and the use of SI units.
- summarize how errors and uncertainty in measurements occur and are reported using significant figures.
- use dimensional analysis effectively.
- determine the density of substances and use density as a conversion factor.

## CONNECTING TO THE LEARNING OBJECTIVES

Your goal is to learn the material in these chapters and our goal is to help you. Use these suggestions to help you to understand the concepts and to master the learning objectives of this chapter.

1. Develop an outline of the chapter, using the section headers (numbered in blue), the sub-headers in red, and the sub-sub-headers in blue. Include the figures and tables you consider important in your outline and include the words highlighted in blue and bold.
2. Construct a chart or table that explains the different measurements including units, conversion factors, and a list of instruments used to make the measurements.
3. After you read the chapter, write three questions about the material presented in the chapter and then write out the answers. If you have a person you study with, then share the questions with each other.

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**FIGURE 1.1 A research chemist working in a modern laboratory.** The controlled conditions in a laboratory permit experiments to yield reproducible results.

# 1.1 Laws and Theories: The Scientific Method

Scientists are curious creatures who want to know what makes the world "tick" (**Figure 1.1**). The approach they take to their work is generally known as the **scientific method** (**Figure 1.2**), which basically boils down to an iterative process of gathering information and formulating explanations.

In the sciences, we usually gather information by performing experiments in laboratories under controlled conditions so the **observations** we make are **reproducible** or can be repeated and the same results obtained. An observation is a statement that accurately describes something we see, hear, taste, feel, or smell. Observations made while performing experiments are referred to as **data**. In order for the observations to be considered valid, they have to be reproducible.

Data gathered during an experiment often lead us to draw conclusions. A **conclusion** is a statement that's based on what we think about a series of observations. For example, consider the following statements about the fermentation of grape juice to make wine:

1. Before fermentation, grape juice is very sweet and contains no alcohol.
2. After fermentation, the grape juice is no longer as sweet and it contains a great deal of alcohol.
3. In fermentation, sugar is converted into alcohol.

Statements 1 and 2 are observations because they describe properties of the grape juice that can be tasted and smelled. Statement 3 is a conclusion because it *interprets* the observations that are available.

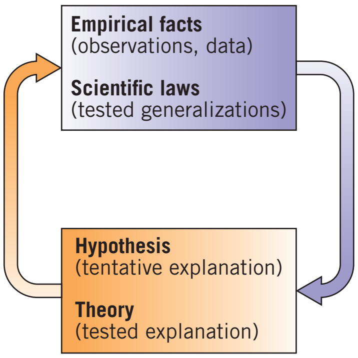


**FIGURE 1.2 The scientific method is iterative** where observations suggest explanations, which suggest new experiments, which suggest new explanations, and so on.

## Experimental Observations and Scientific Laws

One of the goals of science is to organize facts so that relationships or generalizations among the data can be established. For example, if we study the behavior of gases, such as air, we discover that the volume of a gas depends on a number of physical properties, including the amount of the gas, its temperature, and its pressure. The observations relating these physical properties are our data.

By studying data obtained from many different experiments performed at different temperatures and pressures, we observe that when the temperature of a sample of gas is held constant, compressing the gas into half of its original volume causes the pressure of the gas to double. If we were to repeat our experiments many times with different samples of many different gases, we would find that this generalization is applicable to all of them. Such a broad generalization, based on the results of many experiments, is called a **law** or **scientific law**, which is often expressed in the form of mathematical equations. For example, if we represent the pressure of a gas by the symbol $P$ and its volume by $V$, the inverse relationship between pressure and volume can be written as

$$P = \frac{C}{V}$$

where $C$ is a proportionality constant. (We will discuss gases and the laws relating to them in greater detail in Chapter 10.)

**NOTE** The pressure of the gas is inversely proportional to its volume, so the smaller the volume, the larger the pressure.

## Hypotheses and Theories: Models of Nature

As useful as they may be, laws only state what happens; they do not provide explanations. *Why*, for example, *are gases so easily compressed to a smaller volume?* Or, *what must gases be like at the most basic, elementary, level for them to behave as they do?* Answering such questions when they first arise is no simple task and requires much speculation. But over time scientists build mental pictures, called **theoretical models**, that enable them to explain observed laws.

In the development of a theoretical model, researchers form tentative explanations called **hypotheses**. They then perform experiments that test predictions derived from the model. Sometimes the results show that the model is wrong. When this happens, the model must be abandoned or modified to account for the new data. Eventually, if the model survives repeated testing, it achieves the status of a theory. A **theory** is a *well-tested explanation of the behavior of nature*—for example, the atomic theory explains the composition of matter and the changes that matter can undergo. Keep in mind, however, that it is impossible to perform every test that might show a theory to be wrong, so we can never prove *absolutely* that a theory is correct. Furthermore, instruments improve, and scientists are able to collect more and better data. Sometimes this additional data suggests that a theory needs a change, which may be a slight alteration or a complete revision of the theory.

Luck and ingenuity sometimes play an important role in the process of science. For example, in 2004, Andre Geim and Konstantin Novoselov used Scotch tape and a lump of graphite, a form of carbon, to extract a single layer of carbon atoms from the graphite, or *graphene*, **Figure 1.3**. This new extraction technique allowed others to discover graphene's properties, such as being incredibly strong for being only one atom thick and being one of the best conductors of electricity known. Yet, had it not been for Geim and Novoselov's creativity, application of the scientific method to these further discoveries might not have been made.

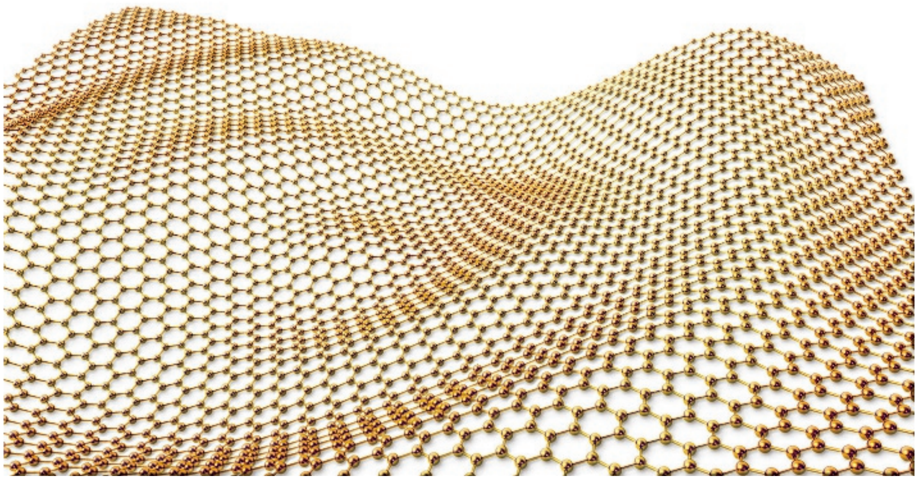

**FIGURE 1.3 Graphene.** One layer of carbon atoms forms an incredibly strong and flexible substance.

As a final note, some of the most spectacular and dramatic changes in science occur when major theories are proved to be wrong. While this happens only rarely, when it does occur, scientists are sent scrambling to develop new theories, and exciting new frontiers are opened.

## 1.1 On the Cutting Edge

### COVID-19

The beginning of the COVID-19 pandemic is an example of how science evolves, and in this case evolved rapidly over a few months. Starting on January 9, 2020, the World Health Organization (WHO) noted that pneumonia-like cases in Wuhan, China, could stem from a coronavirus. At the end of January 2020, the WHO declared a public health emergency, with 9800 cases and more than 200 deaths worldwide, and Wuhan was under a strict quarantine. In February, the US Center for Disease Control (CDC) warned that the outbreak met two of the three conditions for a pandemic: illness resulting in death and sustained person-to-person spread, with the third condition being worldwide spread. On March 11, 2020, the WHO declared COVID-19 a pandemic. By the end of September 2020, over 20 million people were infected and more than 1 million people died from COVID-19 with only 11 countries worldwide reporting no cases.

The science of COVID-19 had to keep pace with this rapidly spreading disease. First, tests had to be developed that could detect the disease. The initial tests took up to 10 days to return a result. Within three months, scientists had developed a test that could indicate infection within an hour. Treatments needed to be established, and the clinical trials required to demonstrate the efficacy of the proposed treatments were held. Finally, a vaccine had to be invented to provide protection for all.

Some of these treatments were proven to be ineffective. For example, there were some initial results that showed that hydroxychloroquine might be effective at treating the disease. However, as more studies were done and more data was collected, the scientists agreed that it was not a possible defense against COVID-19.

On the other hand, the use of masks to protect the user and the people around them, while initially dismissed, has been shown to effectively stop the spread of the disease. This recommendation changed when more was learned about the role that respiratory droplets played in the spread of the virus and this information was paired with numerous clinical and laboratory studies that showed that masks reduce the spray of droplets when worn over the nose and mouth. Again, as the studies were done and the data was collected, science had to change and the recommendations changed.

**PRACTICE EXERCISE 1.1**[1]

What is the process of the scientific method? (*Hint:* Use Figure 1.2.)

## The Atomic Theory as a Model of Nature

Virtually every scientist would agree that one of the most significant theoretical models of nature ever formulated is the atomic theory, which was discussed in some detail in Section 0.4. According to this theory, all chemical substances are composed of tiny submicroscopic particles called **atoms**, which combine in various ways to form all the complex materials we find in the **macroscopic**, visible world around us. This theory forms the foundation for the way scientists think about nature. The atomic theory and how it enables us to explain chemical facts forms the central theme of this book as well.

**NOTE** *Macroscopic* commonly refers to physical objects that are measurable and observable by the naked eye.

# 1.2 Matter and Its Classifications

As mentioned in Section 1.1, one goal of science is to be able to relate things we observe around us and in the laboratory to the way individual atoms and their combinations behave at the submicroscopic level. To begin this discussion, we need to study how chemistry views the macroscopic world.

[1]Answers to the Practice Exercises are found in Appendix B at the back of the book.

## Matter Defined

Chemistry is concerned with the properties and transformations of matter. **Matter** is anything that occupies space and has mass. It is the stuff our universe is made of. All of the objects around us, from rocks to pizza to people, are examples of matter.

Notice that our definition of matter uses the term *mass* rather than *weight*. The words mass and weight are often used interchangeably even though they refer to different things. **Mass** *refers to how much matter there is in a given object, whereas* **weight** *refers to the force with which the object is attracted to another object by gravity*. For example, a golf ball contains a certain amount of matter and has a certain mass, which is the same regardless of the golf ball's location. However, the *weight* of a golf ball on earth is about six times greater than it would be on the moon because the gravitational attraction of the earth is six times that of the moon. Because mass, or amount of matter, does not vary from place to place, we use mass rather than weight when we specify the amount of matter in an object. We measure mass with a balance; Section 1.4 covers this instrument in detail.

## Elements

Matter can be classified into different ways, and one way to classify matter is to consider what it is made of and then to divide it into *elements*, described here, and *compounds* described next. Before we discuss elements, we will define the nature of chemical reactions. Chemistry is especially concerned with **chemical reactions**, which are transformations that alter the chemical compositions of substances. An important type of chemical reaction is **decomposition**, in which one substance is changed into two or more other substances. For example, if we pass electricity through molten (melted) sodium chloride (salt), the silvery metal sodium and the pale green gas chlorine are formed. This change has decomposed sodium chloride into two simpler substances. No matter how we try, however, sodium and chlorine cannot be decomposed further by chemical reactions into still simpler substances that can be stored and studied.

In chemistry, *substances that cannot be decomposed into simpler materials by chemical reactions are called* **elements**. Sodium and chlorine are two examples. Others you may be familiar with include iron, aluminum, sulfur, and carbon (as in graphite, diamonds, and graphene). Some elements are gases at room temperature. Examples include oxygen, nitrogen, hydrogen, chlorine, and helium. Elements are the simplest forms of matter that chemists work with directly. All more complex substances are composed of elements in various combinations. **Figure 1.4** shows some samples of elements that occur uncombined with other elements in nature.

## Chemical Symbols for Elements

So far, scientists have discovered 90 naturally occurring elements and have made 28 more, for a total of 118. Each element is assigned a unique **chemical symbol**, which is used as an

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a. b. c.

**FIGURE 1.4** **Some elements that occur naturally.** **a.** Sulfur. **b.** Gold. **c.** Carbon in the form of diamonds. (Coal is also made up mostly of carbon.)

**TABLE 1.1 Elements That Have Symbols Derived from Their Latin Names**

| Element | Symbol | Latin Name |
|---|---|---|
| Sodium | Na | natrium |
| Potassium | K | kalium |
| Iron | Fe | ferrum |
| Copper | Cu | cuprum |
| Silver | Ag | argentum |
| Gold | Au | aurum |
| Mercury | Hg | hydrargyrum |
| Antimony | Sb | stibium |
| Tin | Sn | stannum |
| Lead | Pb | plumbum |

abbreviation for the name of the element. The names and chemical symbols of the elements are given immediately following the front cover of this text. In most cases, an element's chemical symbol is formed from one or two letters of its English name. For instance, the symbol for carbon is C, for bromine it is Br, and for silicon it is Si. For some elements, the symbols are derived from their non-English names. Those that come from their Latin names are listed in **Table 1.1.** The symbol for tungsten (W) comes from *wolfram*, the German name of the element. Regardless of the origin of the symbol, the first letter is *always* capitalized and the second letter, if there is one, is *always* lowercase.

Chemical symbols are also used to stand for atoms of elements when we write *chemical formulas* such as $H_2O$ (water) and $CO_2$ (carbon dioxide). We will have a lot more to say about formulas later.

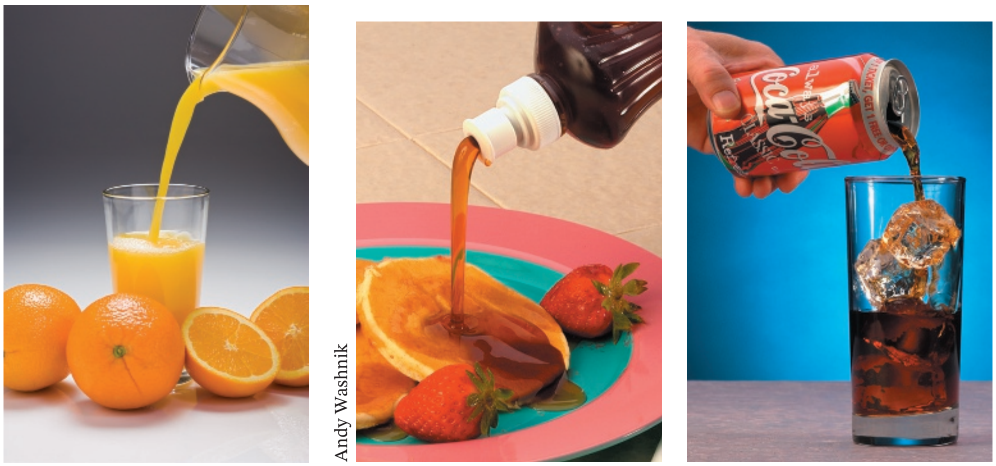
## Compounds

By means of chemical reactions, elements combine in various specific proportions to give all of the more complex substances in nature. Thus, hydrogen and oxygen combine to form water ($H_2O$), and sodium and chlorine combine to form sodium chloride (NaCl, common table salt). Water and sodium chloride are examples of compounds. A compound is a substance formed from two or more different elements in which the elements are always combined in the same, fixed (i.e., constant) proportions by mass. For example, if any sample of pure water is decomposed, the mass of oxygen obtained is always eight times the mass of hydrogen.

## Mixtures

Elements and compounds are examples of **pure substances**.[2] The **composition** of a pure substance is always the same, regardless of its source. Pure substances are rare, however. Usually, we encounter mixtures of compounds or elements. Unlike elements and compounds, **mixtures** *can have variable compositions*. For example, **Figure 1.5** shows three mixtures that contain sugar. They have different degrees of sweetness because the amount of sugar in a given amount varies from one to the other.

Mixtures can be either homogeneous or heterogeneous. A **homogeneous mixture** has the same properties throughout the sample. An example is a thoroughly stirred mixture of sugar in



Juan Silva/Stockbyte/Getty Images
a.
Andy Washnik
b.
c.
Andy Washnik

**FIGURE 1.5 Mixtures can have variable compositions.** Orange juice, pancake syrup, and Coca-Cola are mixtures that contain sugar. The amount of sugar varies from one to another because the composition can vary from one mixture to another.

[2] We have used the term *substance* rather loosely until now. Strictly speaking, **substance** really means *pure substance*. Each unique chemical element and compound is a *substance*; a *mixture* consists of two or more substances.

water. We call such a homogeneous mixture a **solution**. Solutions need not be liquids, just homogeneous. For example, the alloy used in the U.S. 5, 10, and 25 cent coins are solid solutions of copper and nickel, and clean air is a gaseous solution of oxygen, nitrogen, and a number of other gases.

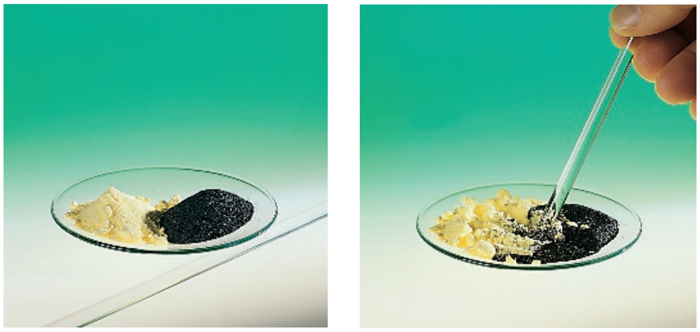
A **heterogeneous mixture** consists of two or more regions, called **phases**, that differ in properties. A mixture of olive oil and vinegar in a salad dressing, for example, is a two-phase mixture in which the oil floats on the vinegar as a separate layer (**Figure 1.6**). The U.S. one cent coin (the penny) is mostly zinc coated with copper. The phases in a mixture don't have to be chemically different substances like oil and vinegar, however. A mixture of ice and liquid water is a two-phase heterogeneous mixture in which the phases have the same chemical composition but occur in different *physical states* (a term we discuss further in Section 1.3).

Andy Washnik

**FIGURE 1.6 A heterogeneous mixture.** The salad dressing contains vinegar and vegetable oil (plus assorted other flavorings). Vinegar and oil do not dissolve in each other, and they form two layers. The mixture is heterogeneous because each of the separate phases (oil, vinegar, and other solids) has its own set of properties that differ from the properties of the other phases.

## Physical and Chemical Changes

The process we use to create a mixture involves a **physical change**, because no new chemical substances form. Shown in **Figure 1.7**, powdered samples of the elements iron and sulfur are simply dumped together and stirred. The mixture forms, but both elements retain their original properties. To separate the mixture, we could similarly use just physical changes. For example, we could remove the iron by stirring the mixture with a magnet—a physical operation. The iron powder sticks to the magnet as we pull it out, leaving the sulfur behind (**Figure 1.8**). The mixture also could be separated by treating it with a liquid called carbon disulfide, which is able to dissolve the sulfur but not the iron. Filtering the sulfur solution from the solid iron, followed by evaporation of the liquid carbon disulfide from the sulfur solution, gives the original components, iron and sulfur, separated from each other.



Michael Watson

a.

Michael Watson

b.

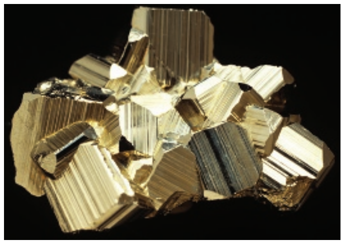
**FIGURE 1.7 Formation of a mixture of iron and sulfur.** **a.** Samples of powdered sulfur and powdered iron. **b.** A mixture of sulfur and iron is made by stirring the two powders together.

Michael Watson

**FIGURE 1.8 Formation of a mixture does not change chemical composition.** Here we see that forming the mixture has not changed the iron and sulfur into a compound of these two elements. The mixture can be separated by pulling the iron out with a magnet. Making a mixture involves a physical change.

The formation of a compound involves a **chemical change** (a chemical reaction) because the chemical makeup of the substances involved are changed. Iron and sulfur, for example, combine to form a compound often called "fool's gold" because it glitters like real gold (**Figure 1.9**). In this compound the elements no longer have the same properties they had before they combined, and they cannot be separated by physical means. The decomposition of fool's gold into iron and sulfur is also a chemical reaction.

The relationships among elements, compounds, and mixtures are shown in **Figure 1.10**.



Charles D. Winters/Science Source

**FIGURE 1.9 "Fools gold."** The mineral pyrite (also called iron pyrite) is a compound of iron and sulfur. When the compound forms, the properties of iron and sulfur disappear and are replaced by the properties of the compound. Pyrite has an appearance that caused some miners to mistake it for real gold.

**PRACTICE EXERCISE 1.2**

We are surrounded by elements, compounds, and mixtures. For example, aluminum foil is pure aluminum, sand is the compound silicon dioxide, and crayons are mixtures of wax and pigments. Classify the following as an element, compound, or mixture. If it is a mixture, decide if it is a homogeneous or heterogeneous mixture: **(a)** lead, **(b)** black coffee, **(c)** sugar, **(d)** lasagna, **(e)** phosphorus. (*Hint:* What are elements, compounds, and mixtures?)

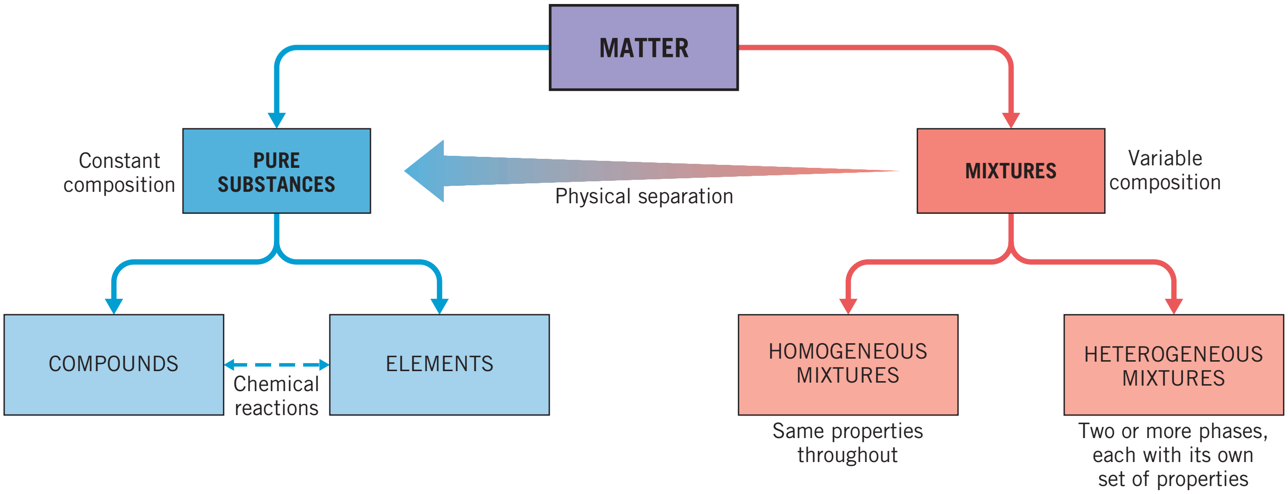



**FIGURE 1.10** **Classification of matter.** This diagram demonstrates the relationships between the different classifications of matter.

## 1.3 Physical and Chemical Properties

In chemistry we use the **properties** (characteristics) of substances to identify them and to distinguish one from another. To organize our thinking, we can classify properties into two different types, physical and chemical, which separate properties by whether the act of observing the property changes the substance, or we categorize properties into extensive or intensive properties, which separates properties by how the property relates to the amount of substance that is being observed. These different types of properties, physical and chemical, or intensive and extensive, are different types of classifications and have different uses in describing matter.

### Physical Properties

The first way to classify properties is whether or not the chemical composition of an object is changed by the act of observing the property. A **physical property** is *one that can be observed without changing the chemical makeup of a substance*. For example, a physical property of gold is that it is yellow. The act of observing this property (color) doesn't change the chemical makeup of the gold. Neither does observing that gold conducts electricity, so color and electrical conductivity are physical properties.

Sometimes, observing a physical property does lead to a physical change. To measure the melting point of ice, for example, we observe the temperature at which the solid begins to melt (**Figure 1.11**). This is a physical change because it does not lead to a change in chemical composition; both ice and liquid water are composed of water molecules.

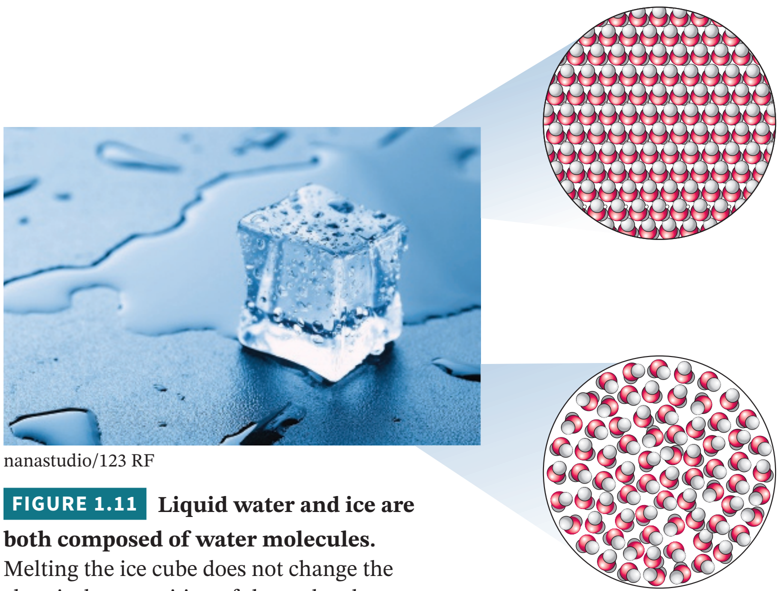


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**FIGURE 1.11** **Liquid water and ice are both composed of water molecules.** Melting the ice cube does not change the chemical composition of the molecules.

### States of Matter

When we observe ice melting, we see two states of matter, liquid and solid. Water can also exist as a gas that we call water vapor or steam. Although ice, liquid water, and steam have quite different appearances and physical properties, they are just different forms of water. **Solid**, **liquid**, and **gas** are the most common states of matter. As with water, most substances are able to exist in all three of these states, and the state we observe generally depends on the temperature. The properties of solids, liquids, and gases relate to the different ways the particles on the atomic scale are organized (**Figure 1.12**), and a change from one state to another is a physical change.

## Chemical Properties

A **chemical property** *describes how a substance undergoes a chemical change (chemical reaction)*. When a chemical reaction takes place, the chemicals interact to form entirely *different* substances with different chemical and physical properties. As previously described, pyrite is formed by the chemical reaction of sulfur and iron, and the product does not resemble sulfur or iron. It's a golden solid that is not attracted by a magnet.

The ability of iron to react with sulfur to form pyrite is a chemical property of iron. When we observe this property, we conduct a reaction, so after we've made the observation, we no longer have the same substances as before. In describing a chemical property, we usually refer to a chemical reaction.

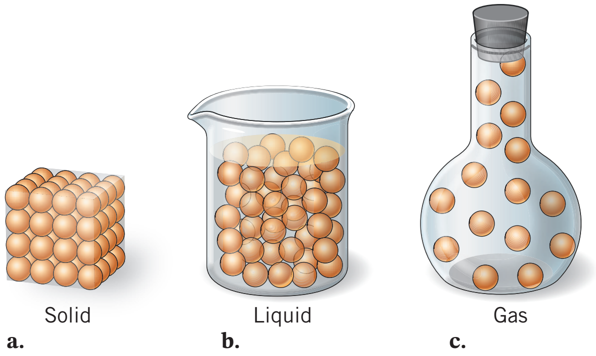



**FIGURE 1.12 Solid, liquid, and gaseous states viewed using the atomic model of matter. a.** In a solid, the particles are tightly packed and cannot move easily. **b.** In a liquid, the particles are still close together but can readily move past one another. **c.** In a gas, particles are far apart with much empty space between them.

## Intensive and Extensive Properties

Another way of classifying a property is according to whether or not it depends on the size of the sample. For example, two different pieces of gold can have different volumes, but both have the same characteristic shiny yellow color and both will begin to melt at the same temperature. Color and melting point (and boiling point, too) are examples of **intensive properties**—*properties that are independent of sample size*. Volume, on the other hand, is an **extensive property**—a *property that depends on sample size*. Mass is another extensive property.

A job chemists often perform is *chemical analysis*. They're asked, "What is a particular sample composed of?" To answer such a question, the chemist relies on the properties of the chemicals that make up the sample. For identification purposes, intensive properties are more useful than extensive ones because every sample of a given substance exhibits the same set of intensive properties.

Color, freezing point, and boiling point are examples of intensive physical properties that can help us identify substances. Chemical properties are also intensive properties and also can be used for identification. For example, gold miners were able to distinguish between real gold and fool's gold, or pyrite, (Figure 1.9) by heating the material in a flame. Nothing happens to the gold, but the pyrite sputters, smokes, and releases bad-smelling fumes because of its ability, when heated, to react chemically with oxygen in the air.

### PRACTICE EXERCISE 1.3

Each of the following can be classified as a physical or chemical change. Identify which type of change is given and explain your choice.

**(a)** Sugar caramelizes.

**(b)** Iron wire glows red when an electric current is passed through it.

**(c)** Liquid nitrogen boils at −196 °C.

**(d)** Tea is made sweeter by adding sugar.

### PRACTICE EXERCISE 1.4

Each of these properties can be classified as intensive or extensive. Identify which type of property is given and explain your choice.

**(a)** The color of a blue paint chip.

**(b)** The amount of paint needed to paint a room.

**(c)** The thickness of the paint when it is well stirred.

**(d)** The weight of the paint left in the can after the room is painted.

# 1.4 Measurement of Physical and Chemical Properties

## Qualitative and Quantitative Observations

Earlier you learned that an important step in the scientific method is observation. In general, observations fall into two categories, qualitative and quantitative. **Qualitative observations**, such as the color of a chemical or that a mixture becomes hot when a reaction occurs, do not involve numerical information. **Quantitative observations** are those **measurements** that do yield numerical data. You make such observations in everyday life, for example, when you glance at your watch or step onto a bathroom scale. In chemistry, we make various measurements that aid us in describing both chemical and physical properties.

## Measurements Include Units

Measurements involve numbers, but they differ from the numbers used in mathematics in two crucial ways.

First, measurements always involve a comparison to a standard that has the units, and therefore need a *unit of measurement*. When you say that a person is six feet tall, you're really saying that the person is six times taller than a reference object that is 1 foot in length, where foot is an example of a **unit of measurement**. *Both the number and the unit are essential parts of the measurement*, because the unit gives a sense of size. For example, if you were told that the distance between two points is 25, you would naturally ask "25 what?" The distance could be 25 inches, 25 feet, 25 miles, or 25 of any other unit that's used to express distance. This demonstrates that a number without a unit is really meaningless.

The second important difference is that measurements always involve uncertainty; they are *inexact*. The act of measurement involves an estimation of one sort or another, and both the observer and the instruments used to make the measurement have inherent physical limitations. As a result, measurements always include some uncertainty, which can be minimized but never entirely eliminated. We will say more about this topic in Section 1.5.

## International System of Units (SI Units)

A standard system of units is essential if measurements are to be made consistently. In the sciences, metric-based units are used. The advantage of working with metric units is that converting to larger or smaller values can be done simply by moving a decimal point, because metric units are related to each other by simple multiples of ten.

The original metric system has been simplified and now is called the **International System of Units**, abbreviated **SI** from the French name, *Le Système International d'Unités.* The SI is the dominant system of units in science and engineering, although there is still some usage of older, outdated, units.

The SI has as its foundation a set of **base units** for seven measured quantities (**Table 1.2**). For now, we will focus on the base units for length, mass, time, and temperature. We will discuss the unit for amount of substance, the mole, at length in Chapter 3. The unit for electrical current,

TOOLS
**Base SI Units**

**TABLE 1.2 The SI Base Units**

| Measurement | Unit | Symbol |
|---|---|---|
| Length | meter | m |
| Mass | kilogram | kg |
| Time | second | s |
| Electric current | ampere | A |
| Temperature | kelvin | K |
| Amount of substance | mole | mol |
| Luminous intensity | candela | cd |

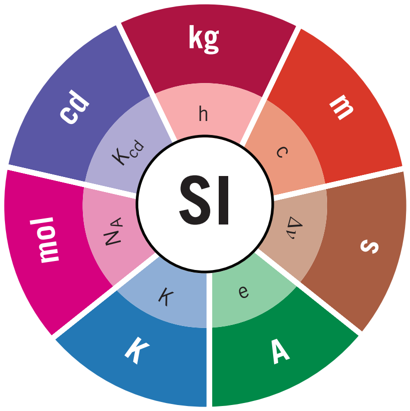
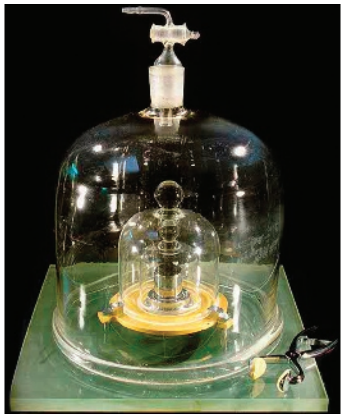



**FIGURE 1.13** **The SI logo produced by the Bureau International des Poids et Mesures** showing the seven SI base units and the seven defining constants.

the ampere, will be discussed briefly when we study electrochemistry in Chapter 19. The unit for luminous intensity, the candela, will not be discussed in this book.

In 2019, the International Bureau of Weights and Measures (Bureau International des Poids et Mesures, BIPM), an intergovernmental organization that regulates the SI, defined the magnitude of the seven base units in terms of seven constants (**Figure 1.13**). Each unit can now be expressed through a fundamental constant in nature, such as the speed of light, $c$, or the charge of a proton, $e$. These constants were fixed to their best estimates in 2019 and can be used to define the base units; before 2019, the constants were just very precisely measured quantities. For instance, the meter is now defined as exactly the distance light travels in a vacuum in 1/299792458 of a second. This change has had the greatest impact on the kilogram, which used to be defined by an object made by human hands, a carefully preserved platinum–iridium alloy cylinder stored at the International Bureau of Weights and Measures in France (**Figure 1.14**).



NIST

**FIGURE 1.14** **The international standard kilogram.** The pre-2019 SI standard for mass is made of a platinum-iridium alloy and is protected under two bell jars, as shown, at the International Bureau of Weights and Measures in France.

## 1.2 On the Cutting Edge

### Redefining Measurements

A unit is an example of a quantity that can be used as a reference for a measurement. The SI units have all be defined in terms of seven fundamental constants whose values have been fixed. By using either the constant itself or by combining the constants, all of the units can be determined. The units used in the defining constants are hertz (Hz), meter (m), second (s), Kelvin (K), mole (mol), joule (J), coulomb (C), lumen (lm), and watt (W), as shown in **Table 1**.

**TABLE 1** **The Units We Need**

| Measurement | Unit | Symbol | Equivalency |
|---|---|---|---|
| Frequency | hertz | Hz | $s^{-1}$ or 1/s or $\frac{1}{s}$ |
| Length | meter | m | Base unit |
| Time | second | s | Base unit |
| Temperature | Kelvin | K | Base unit |
| Amount of material | mole | mol | Base unit |
| Energy | joule | J | $kg\ m^2\ s^{-2}$ or $\frac{kg\ m^2}{s^2}$ |
| Electrical current | ampere | A | Base unit |
| Luminous intensity | candela | cd | Base unit |
| Luminous intensity | lumen | lm | cd sr[1] |
| Solid angle | steradian | sr | See Figure 1 |
| Power | watt | W | $kg\ m^2\ s^{-3}$ or $\frac{kg\ m^2}{s^3}$ |

[1]See Figure 1 for the steradian (sr).

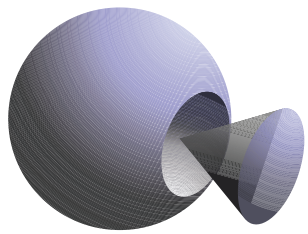


**FIGURE 1** **A steradian (sr)** is a solid angle equal to the angle at the center of a sphere subtended by a part of the surface equal in area to the square of the radius. There are $4\pi$ steradians in the surface area of a sphere.

The seven defining constants for the SI base units are presented in Table 2.

**Table 2 The Seven Fundamental Constants Used to Define the SI Base Units**

| Defining Constant | Symbol | Numerical Value and Unit |
|---|---|---|
| Hyperfine transition frequency of Cs-133[1] | $\Delta\nu_{Cs}$ | 9,192,631,770 Hz |
| Speed of light in vacuum | $c$ | 299,792,458 m/s |
| Planck constant | $h$ | $6.62607015 \times 10^{-34}$ J s |
| Elementary charge | $e$ | $1.602176634 \times 10^{-19}$ C |
| Boltzmann constant | $K_B$ | $1.380649 \times 10^{-23}$ J K$^{-1}$ |
| Avogadro constant | $N_A$ | $6.02214076 \times 10^{23}$ mol$^{-1}$ |
| Luminous efficacy[2] | $K_{cd}$ | 683 lm W$^{-1}$ |

[1]The unperturbed ground state hyperfine transition frequency of the cesium-133 atom $\Delta\nu_{Cs}$ is 9,192,631,770 Hz (9,192,631,770 s$^{-1}$).

[2]The luminous efficacy of monochromatic radiation of frequency $540 \times 10^{12}$ Hz, $K_{cd}$, is 683 lm/W.

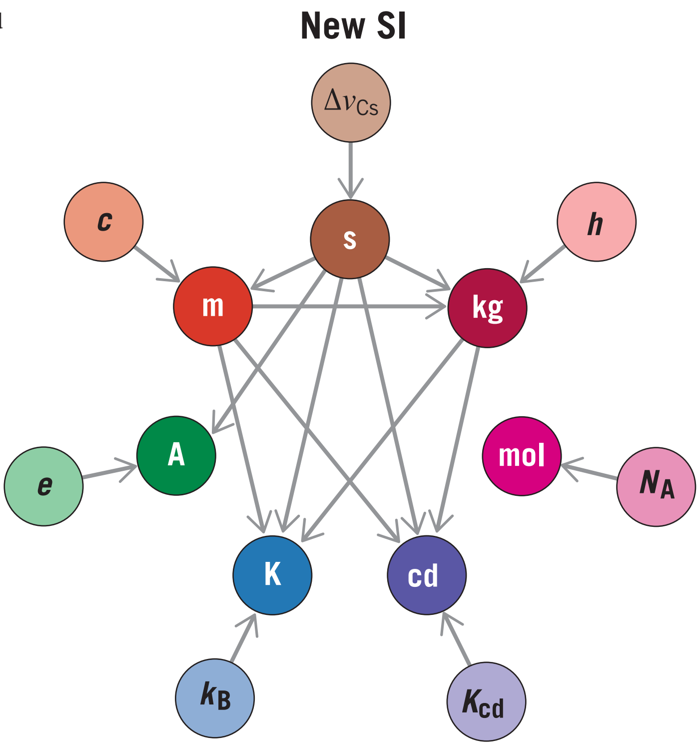



**FIGURE 2 Relationship between SI base units and physical constants.** In the new SI, the base units are defined in terms of the seven physical constants with fixed numerical values and on other base units that are derived from the same set of constants. The diagram shows the relationships between the units and the constants.

Using the fundamental constants, we can now define the seven base SI units (Figure 2 and Table 3).

Finally, one more change that has been standardized is how numbers are written. For large numbers and small numbers less than one with many decimal places, spaces are used to separate the numbers. For example, 9,192,631,770 Hz would be written as 9 192 631 770 Hz and $1.602176634 \times 10^{-19}$ C would be written as $1.602\ 176\ 634 \times 10^{-19}$ C. In this book, we will be using commas if we have a number that is larger than 9999 (e.g., 9,192,631,770 Hz). Furthermore, if a number is less than one and is not written in scientific notation, then a zero always precedes the decimal point (e.g., write 0.393 7 inches, not .3937 inches).

**TABLE 3 The Definition of the Seven SI Base Units[1]**

| Unit | Symbol | Definition[1] |
|---|---|---|
| Second (s) | $t$ | Defined by taking the fixed numerical value of the cesium frequency $\Delta\nu_{Cs}$, the unperturbed ground-state hyperfine transition frequency of the cesium-133 atom, to be 9,192,631,770 when expressed in the unit Hz, which is equal to s$^{-1}$. |
| Meter (m) | $l$, $d$, | Defined by taking the fixed numerical value of the speed of light in vacuum $c$ to be 299,792,458 when expressed in the unit m s$^{-1}$, where the second is defined in terms of $\Delta\nu_{Cs}$. |
| Kilogram (kg) | $m$ | Defined by taking the fixed numerical value of the Planck constant, $h$, to be $6.62607015 \times 10^{-34}$ when expressed in the unit J s, which is equal to kg m$^2$ s$^{-1}$, where the meter and the second are defined in terms of $c$ and $\Delta\nu_{Cs}$. |
| Ampere (A) | $I$ | Defined by taking the fixed numerical value of the elementary charge $e$ to be $1.602176634 \times 10^{-19}$ when expressed in the unit C, which is equal to A s, where the second is defined in terms of $\Delta\nu_{Cs}$. |
| Kelvin (K) | $T$ | Defined by taking the fixed numerical value of the Boltzmann constant $k$ to be $1.380649 \times 10^{-23}$ when expressed in the unit J K$^{-1}$, which is equal to kg m$^2$ s$^{-2}$ K$^{-1}$, where the kilogram, meter, and second are defined in terms of $h$, $c$, and $\Delta\nu_{Cs}$. |
| Mole (mol) | $n$ | The amount of substance of exactly $6.02214076 \times 10^{23}$ elementary entities. This number is the fixed numerical value of the Avogadro constant, $N_A$, when expressed in the unit mol$^{-1}$. |
| Candela (cd) | $I_V$ | Defined by taking the fixed numerical value of the luminous efficacy of monochromatic radiation of frequency $540 \times 10^{12}$ Hz, $K_{cd}$, to be 683 when expressed in the unit lm W$^{-1}$, which is equal to cd sr W$^{-1}$, or cd sr kg$^{-1}$ m$^{-2}$ s$^3$, where the kilogram, meter, and second are defined in terms of $h$, $c$, and $\Delta\nu_{Cs}$. |

[1]National Institute of Standards and Technology, https://www.nist.gov/pml/weights-and-measures/metric-si/si-units, accessed June 2021.

[2]Le Système international d'unités, The International System of Units, 9e édition, 2019.

In scientific measurements, *all* physical quantities will have units that are combinations of the seven base SI units. For example, there is no SI base unit for area, but to calculate area we multiply length by width. Therefore, the *unit* for area is derived by multiplying the *unit* for length by the *unit* for width. Length and width are measurements that have the SI base unit of the **meter (m)**.

$$\text{length} \times \text{width} = \text{area}$$
$$(\text{meter}) \times (\text{meter}) = (\text{meter})^2$$
$$\text{m} \times \text{m} = \text{m}^2$$

The SI **derived unit** for area is therefore $\text{m}^2$ (read as *meters squared* or *square meter*).

In deriving SI units, we employ a very important concept that we will use repeatedly throughout this book when we perform calculations: *Units undergo the same kinds of mathematical operations that numbers do*. We will see how this fact can be used to convert from one unit to another in Section 1.6.

## EXAMPLE 1.1 | Deriving SI Units

*Linear momentum* is the term used to describe an object moving at a constant velocity in a straight line. It is equal to the object's mass times its velocity. What is the SI derived unit for linear momentum?

**Analysis:** To derive a unit for a quantity, first we express it in terms of simpler quantities. We're told that linear momentum is mass times velocity. Therefore, the SI unit for linear momentum will be the SI unit for mass times the SI unit for velocity. Since velocity is not one of the base SI units, we will need to find the SI units for velocity.

**Assembling the Tools:** The tool we need is the list of SI base units in Table 1.2. We will also have to recall that velocity is distance traveled (length) per unit time.

**Solution:** We start by expressing the given information as an equation:

$$\text{mass} \times \text{velocity} = \text{linear momentum}$$

Next we write the same equation with units, kg for mass and m/s for velocity, because units undergo the same mathematical operations as the numbers.

$$\text{kg} \times \text{m/s} = \text{kg m/s or kg m s}^{-1}$$

Our result is that linear momentum has units of kg m/s or kg m $\text{s}^{-1}$.

**Is the Answer Reasonable?** *Before leaving a problem, it is always wise to examine the answer to see whether it makes sense*. In this problem, the check is simple. The derived unit for linear momentum should be the product of units for mass and velocity, and this is obviously true.

**NOTE** Notice that for derived units, scientists usually omit the multiplication sign but always leave a space between units.

### PRACTICE EXERCISE 1.5

The volume of a sphere is given by the formula $V = \frac{4}{3}\pi r^3$, where $r$ is the radius of the sphere. From this equation, determine the SI units for volume. (*Hint: r* is a length, so it must have a length unit.)

### PRACTICE EXERCISE 1.6

When you "step hard on the gas" in a car you feel an invisible force pushing you back in your seat. This force, $F$, equals the product of your mass, $m$, times the acceleration, $a$, of the car. In equation form, this is $F = ma$. Acceleration is the change in velocity, $v$, with time, $t$:

$$a = \frac{\text{change in } v}{\text{change in } t}$$

Therefore, the units of acceleration are those of velocity divided by time. What is the SI derived unit for force expressed in SI base units?


Michael Watson

**FIGURE 1.15** **Metric units are commonplace on consumer products.**

**TABLE 1.3 Some Non-SI Metric Units Commonly Used in Chemistry**

| Measurement | Unit | Symbol | Value in SI Units |
|---|---|---|---|
| Length | angstrom | Å | 1 Å = 0.1 nm = $10^{-10}$ m |
| Mass | atomic mass unit | u (amu) | 1 u = $1.66054 \times 10^{-27}$ kg (rounded to six digits) |
| | metric ton | t | 1 t = $10^3$ kg |
| Time | minute | min. | 1 min. = 50 s |
| | hour | h | 1 h = 60 min. = 3600 s |
| Temperature | degree Celsius | °C | $T_K = t_{°C} + 273.15$ |
| Volume | liter | L | 1 L = 1000 $cm^3$ |

**TABLE 1.4 Some Useful Conversions**

| Measurement | English Unit | English/SI Equality[a] |
|---|---|---|
| Length | inch | 1 in. = 2.54 cm |
| | yard | 1 yd = 0.9144 m |
| | mile | 1 mi = 1.609 km |
| Mass | pound | 1 lb = 453.6 g |
| | ounce (mass) | 1 oz = 28.35 g |
| Volume | gallon | 1 gal = 3.785 L |
| | quart | 1 qt = 946.4 mL |
| | ounce (fluid) | 1 oz = 29.6 mL |

[a]These equalities allow us to convert English to metric or metric to English units.

## Other Unit Systems

Some older metric units that are not part of the SI system are still used in the laboratory and in the scientific literature. Some of these units are listed in **Table 1.3**; others will be introduced as needed in upcoming chapters.

The United States is one of the only nations still using the **English system** of units, which measures distance in inches, feet, and miles; volume in ounces, quarts, and gallons; and mass in ounces and pounds. However, many of the English units are defined with reference to base SI units. Beverages, food packages, tools, and machine parts are often labeled in metric units (**Figure 1.15**). Common conversions between the English system and the SI are given in **Table 1.4** and immediately preceding the back cover of this text.[3]

## Decimal Multipliers

Sometimes the basic units are either too large or too small to be used conveniently. For example, the meter is inconvenient for expressing the size of very small things such as bacteria. The SI solves this problem by forming larger or smaller units by applying **decimal multipliers** to the base units. **Table 1.5** lists the decimal multipliers and the prefixes used to identify them. These were set by the International Bureau of Weights and Measures in 2019. Those listed in red type are the ones most commonly encountered in chemistry.

**NOTE** A modified base unit can be conveniently converted to the base unit by removing the prefix and inserting $\times 10^x$.

$25.2 \text{ pm} = 25.2 \times 10^{-12} \text{ m}$

When the name of a unit is preceded by one of these prefixes, the size of the unit is modified by the corresponding decimal multiplier. For instance, the prefix *kilo* indicates a multiplying factor of $10^3$, or 1000. Therefore, a *kilo*meter is a unit of length equal to

[3]Originally, these conversions were established by measurement. For example, if a metric ruler is used to measure the length of an inch, it is found that 1 in. equals 2.54 cm. Later, to avoid confusion about the accuracy of such measurements, it was agreed that these relationships would be taken to be exact. For instance, 1 in. is now defined as *exactly* 2.54 cm. Exact relationships also exist for the other quantities, but for simplicity many have been rounded off. For example, 1 lb = 453.59237 g, *exactly*.

**TABLE 1.5** SI Prefixes—Their Meanings and Values[a]

| Prefix | Meaning | Symbol | Prefix Value[b] (numerical) | Prefix Value[b] (power of ten) |
|---|---|---|---|---|
| yotta | | Y | | $10^{24}$ |
| zetta | | Z | | $10^{21}$ |
| exa | | E | | $10^{18}$ |
| peta | | P | | $10^{15}$ |
| tera | | T | | $10^{12}$ |
| giga | billions of | G | 1000000000 | $10^{9}$ |
| mega | millions of | M | 1000000 | $10^{6}$ |
| kilo | thousands of | k | 1000 | $10^{3}$ |
| hecto | | h | | $10^{2}$ |
| deka | | da | | $10^{1}$ |
| deci | tenths of | d | 0.1 | $10^{-1}$ |
| centi | hundredths of | c | 0.01 | $10^{-2}$ |
| milli | thousandths of | m | 0.001 | $10^{-3}$ |
| micro | millionths of | m | 0.000001 | $10^{-6}$ |
| nano | billionths of | *n* | 0.000000001 | $10^{-9}$ |
| pico | trillionths of | p | 0.000000000001 | $10^{-12}$ |
| femto | | f | | $10^{-15}$ |
| atto | | a | | $10^{-18}$ |
| zepto | | z | | $10^{-21}$ |
| yocto | | y | | $10^{-24}$ |

[a]Prefixes in red type are used most often.
[b]Numbers in these columns can be interchanged with the corresponsding prefix.

**TOOLS**
**SI prefixes**

1000 meters.[4] The symbol for kilometer (km) is formed by applying the symbol meaning kilo (k) as a prefix to the symbol for meter (m). Thus 1 km $=$ 1000 m (or 1 km $= 10^3$ m). Similarly a decimeter (dm) is 1/10th of a meter, so 1 dm $=$ 0.1 m (1 dm $= 10^{-1}$ m).

**NOTE** A prefix can be inserted by adjusting the exponential part of a number to match the definition of a prefix. Then the 10 raised to an exponent is removed and replaced by the prefix.

$2.34 \times 10^{10}$ g $= 23.4 \times 10^{9}$ g $=$ 23.4 Gg

## Laboratory Measurements

Length, volume, mass, and temperature are the most common measurements made in the laboratory.

**TOOLS**
**Units in laboratory measurements**

**Length** The SI base unit for length, the meter (m), is too large for most laboratory purposes. More convenient units are the **centimeter (cm)** and the **millimeter (mm)**. Using Table 1.5 we see that they are related to the meter as follows.

$$1 \text{ cm} = 10^{-2} \text{ m} = 0.01 \text{ m}$$
$$1 \text{ mm} = 10^{-3} \text{ m} = 0.001 \text{ m}$$

[4]In the sciences, powers of 10 are often used to express large and small numbers. The quantity $10^3$ means $10 \times 10 \times 10 = 1000$. Similarly, the quantity $6.5 \times 10^2 = 6.5 \times 10 \times 10 = 650$. Numbers less than 1 have negative exponents when expressed as powers of 10. Thus, the fraction 1/10 is expressed as $10^{-1}$, so the quantity $10^{-3}$ means $1/10 \times 1/10 \times 1/10 = 1/1000$. A value of $6.5 \times 10^{-3} = 6.5 \times 1/10 \times 1/10 \times 1/10 = 0.0065$. Numbers written as $6.5 \times 10^2$ and $6.5 \times 10^{-3}$, with the decimal point between the first and second digit, are said to be expressed in scientific notation.

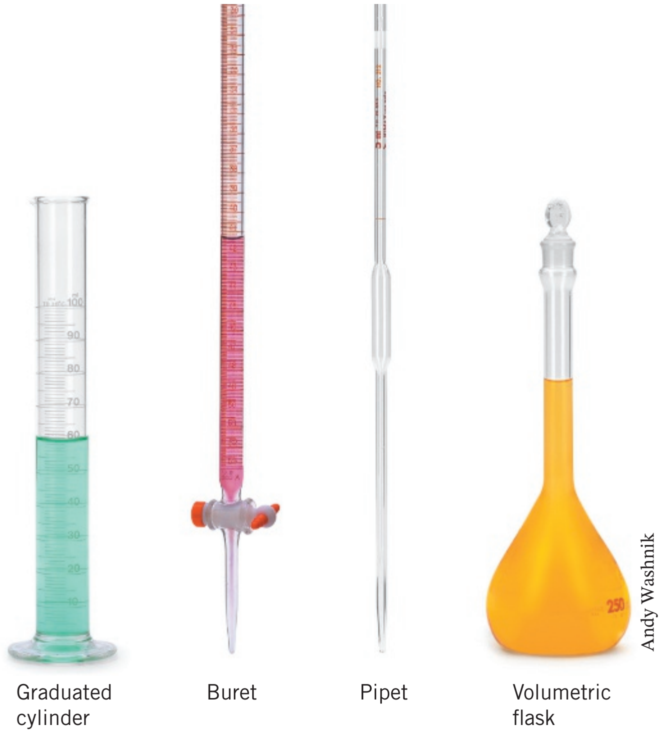



**FIGURE 1.16 Common laboratory glassware used for measuring volumes.** Graduated cylinders are used for measuring volumes to the nearest milliliter. Precise measurements of volumes are made using burets, pipets, and volumetric flasks.

It is also useful to know the relationships

$$1 \text{ cm} = 100 \text{ cm} = 1000 \text{ cm}$$
$$1 \text{ cm} = 10 \text{ mm}$$

**NOTE** An older, non-SI unit called the angstrom (Å) is often used to describe dimensions of atomic- and molecular-sized particles.

$1 \text{ Å} = 0.1 \text{ nm} = 10^{-10} \text{ m}$

## Volume

Volume is a derived unit with dimensions of $(\text{length})^3$. With these dimensions expressed in meters, the derived SI unit for volume is the **cubic meter, $m^3$**.

In chemistry, measurements of volume usually arise when we measure amounts of liquids. The traditional metric unit of volume used for this is the **liter (L)**. In SI terms, a liter is defined as exactly 1 cubic decimeter.

$$\boxed{1 \text{ L} = 1 \text{ dm}^3 \quad (\text{exactly})} \tag{1.1}$$

However, even the liter is too large to conveniently express most volumes measured in the lab. The glassware we normally use, as illustrated in **Figure 1.16**, is marked in **milliliters (mL)**.[5]

$$1 \text{ L} = 1000 \text{ mL}$$

Because 1 dm = 10 cm, then $1 \text{ dm}^3 = 1000 \text{ cm}^3$. Therefore, 1 mL is exactly the same as $1 \text{ cm}^3$.

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

Sometimes you may see $\text{cm}^3$ abbreviated cc for cubic centimeter (especially in medical applications), although the SI frowns on this symbol. **Figure 1.17** compares the cubic meter, liter, and milliliter.

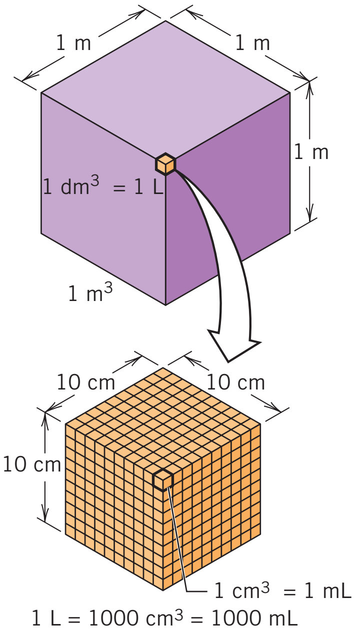


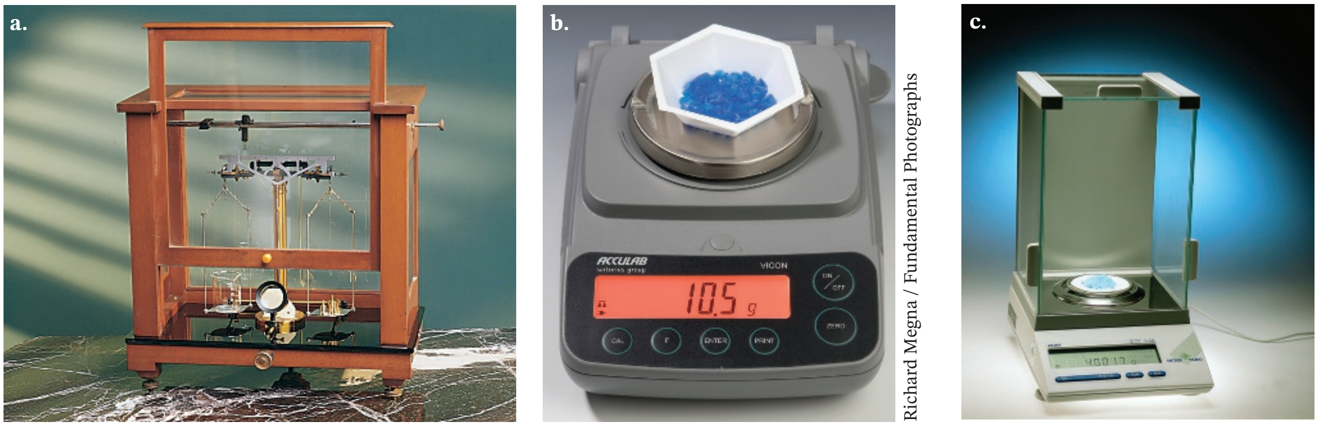

**FIGURE 1.17 Comparing volume units.** A cubic meter ($m^3$) is approximately equal to a cubic yard; 1000 $cm^3$ is approximately a quart, and approximately 30 $cm^3$ is equal to one fluid ounce.

## Mass

In the SI, the base unit for mass is the **kilogram (kg)**, although the **gram (g)** is a more conveniently sized unit for most laboratory measurements. One gram, of course, is 1/1000 of a kilogram (1 kilogram = 1000 g, so 1 g must equal 0.001 kg).

[5]Use of the abbreviations L for liter and mL for milliliter is rather recent. Confusion between the printed letter l and the number 1 prompted the change from l to L and ml to mL. You may encounter the abbreviation ml in other books or on older laboratory glassware.

Mass is measured by comparing the weight of a sample with the weights of known standard masses. (Recall that mass and weight are not the same thing.) The instrument used is called a **balance** (**Figure 1.18**). For the balance in Figure 1.18*a*, we would place our sample on the left pan and then add standard masses to the other. When the weight of the sample and the total weight of the standards are in balance (when they match), their masses are then equal. **Figure 1.19** shows the masses of some common objects in SI units.

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Richard Megna / Fundamental Photographs

Charles D. Winters/Science Source

**FIGURE 1.18** **Typical laboratory balances.** **a.** A traditional two-pan analytical balance capable of measurements to the nearest 0.0001 g. **b.** A modern top-loading balance capable of mass measurements to the nearest 0.1 g. **c.** A modern analytical balance capable of measurements to the nearest 0.0001 g.

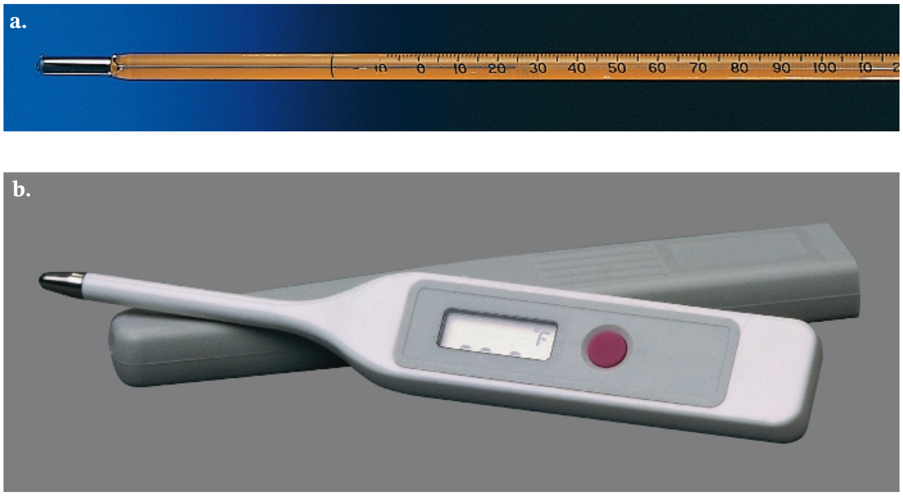

**FIGURE 1.19** **Masses of several common objects.**

## Temperature

Temperature is usually measured with a thermometer (**Figure 1.20**). Thermometers are graduated in *degrees* according to one of two temperature scales. On the **Fahrenheit scale**, water freezes at 32 °F and boils at 212 °F. (See **Figure 1.21**.) On the **Celsius scale**, water freezes at 0 °C and boils at 100 °C, which means there are 100 degree units between the freezing and boiling points of water, while on the Fahrenheit scale, this same temperature range is spanned by 180 degree units. Consequently, five Celsius degrees are the same as nine Fahrenheit degrees. We can use the following equation as a tool to convert between these temperature scales.

$$\boxed{t_F = \left(\frac{9\ °F}{5\ °C}\right)t_C + 32\ °F} \qquad (1.2)$$

**TOOLS**
**Celsius-to-Fahrenheit conversions**

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Lawrence Manning/Corbis/Getty Images

**FIGURE 1.20 Typical laboratory thermometers.** **a.** A traditional mercury thermometer. **b.** A home medical thermometer.

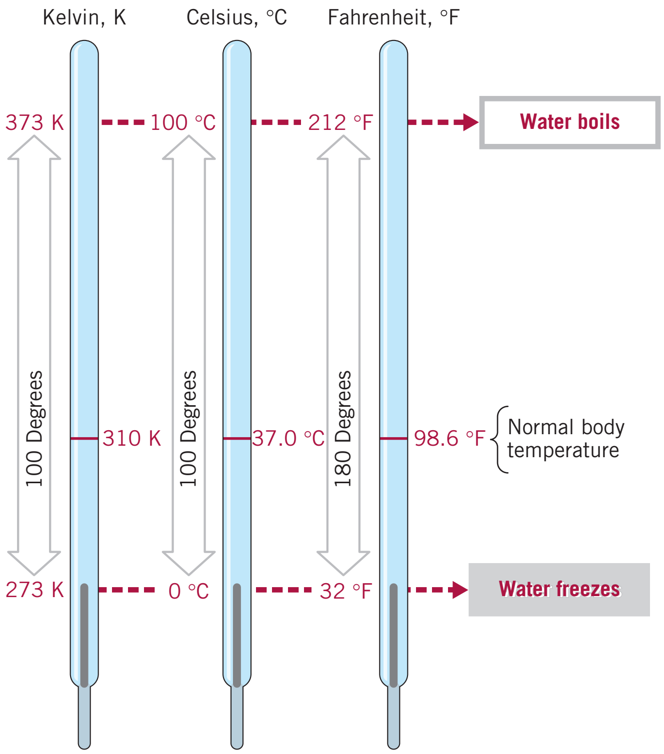



**FIGURE 1.21 Comparison of temperature scales.** Kelvin, Celsius, and Fahrenheit temperature scales.

**NOTE** The name of the temperature scale, the Kelvin scale, is capitalized, but the name of the unit, the kelvin, is not. However, the symbol for the kelvin is a capital K.

In this equation, $t_F$ is the Fahrenheit temperature and $t_C$ is the Celsius temperature. As noted earlier, units behave like numbers in calculations, and we see in Equation 1.2 that °C "cancels out" to leave only °F. The 32 °F is added to account for the fact that the freezing point of water (0 °C) occurs at 32 °F on the Fahrenheit scale. Equation 1.2 can easily be rearranged to permit calculating the Celsius temperature from the Fahrenheit temperature.

$$(t_F - 32\ °F)\left(\frac{5\ °C}{9\ °F}\right) = t_C$$

The SI unit of temperature is the **kelvin (K)**, which is the degree unit on the **Kelvin scale**. Notice that the temperature unit is K, not °K (the degree symbol, °, is omitted). Also notice that the name of the unit, kelvin, is not capitalized.

Figure 1.20 shows how the Kelvin, Celsius, and Fahrenheit temperature scales relate to each other. Notice that the kelvin is *exactly* the same size as the Celsius degree. *The only difference between these two temperature scales is the zero point.* The zero point on the Kelvin scale is called **absolute zero** and corresponds to nature's lowest temperature. It is 273.15 degree units below the zero point on the Celsius scale, which means that 0 °C equals 273.15 K and 0 K equals −273.15 °C. To convert from Celsius to Kelvin temperatures the following equation applies.

**TOOLS** Celsius-to-Kelvin conversions

$$T_K = (t_C + 273.15)\,\cancel{°C}\,\frac{1\ \text{K}}{1\ \cancel{°C}} \tag{1.3}$$

This amounts to simply adding 273.15 to the Celsius temperature to obtain the Kelvin temperature. Often we are given Celsius temperatures rounded to the nearest degree, in which case we round 273.15 to 273. Thus, 25 °C equals (25 + 273) K or 298 K.

**NOTE** We will use a capital *T* to stand for the Kelvin temperature and a lowercase *t* (as in $t_C$) to stand for the Celsius temperature. This conforms to the usage described by the International Bureau of Weights and Measures in Sevres, France, and the National Institute of Standards and Technology in Gaithersburg, Maryland.

### Types of Mathematical Calculations

We will discuss many problems throughout the book that require mathematical calculations. In approaching them you will see that they can be divided into two general types. The type described in Example 1.2 involves applying a mathematical equation in which we have numerical values for all but one of the variables in the equation. To perform the calculation, we solve the equation by substituting known values for the variables until there is only one unknown variable that can be calculated. The second type of calculation that we will encounter, which is described in detail in Section 1.6, is one where we convert one set of units into another set of units using a method called dimensional analysis.

If you follow our approach of *Analysis* and *Assembling the Tools*, the nature of the types of calculations involved should become apparent. As you will see, some problems involve both kinds of calculations.

## EXAMPLE 1.2 | Converting Among Temperature Scales

*Thermal pollution*, the release of large amounts of heat into rivers and other bodies of water, is a serious problem near power plants and can affect the survival of some species of fish. For example, trout will die if the temperature of the water rises above approximately 25 °C. **(a)** What is this temperature in °F? **(b)** Rounded to the nearest whole degree unit, what is this temperature in kelvins?

**Analysis:** Both parts of the problem here deal with temperature conversions. Therefore, we ask, "What tools do we have that relate the temperature scales to each other?"

**Assembling the Tools:** The first tool, Equation 1.2, relates Fahrenheit temperatures to Celsius temperatures and is needed to answer part **(a)**.

$$t_F = \left(\frac{9\ °\text{F}}{5\ °\text{C}}\right)t_C + 32\ °\text{F}$$

Equation 1.3 relates Kelvin temperatures to Celsius temperatures and is needed for part **(b)**.

$$T_K = (t_C + 273.15)\,°\text{C}\,\frac{1\ \text{K}}{1\ °\text{C}}$$

**Solution:**

**Part (a):** We substitute the value of the Celsius temperature (25 °C) for $t_C$ and calculate the answer

$$t_F = \left(\frac{9\ °F}{5\ \cancel{°C}}\right)(25\ \cancel{°C}) + 32\ °F$$
$$= 45\ °F + 32\ °F = 77\ °F$$

Therefore, 25 °C = 77 °F. (Notice that we have canceled the unit °C in the equation above.)

**Part (b):** Once again, we have a simple substitution. Since $t_C = 25$ °C, the Kelvin temperature (rounded) is calculated as

$$T_K = (t_C + 273.15)°C\,\frac{1\ K}{1\ °C}$$
$$= (25 + 273.15)°C\,\frac{1\ K}{1\ °C}$$
$$= 298\ \cancel{°C}\,\frac{1\ K}{1\ \cancel{°C}} = 298\ K$$

Thus, 25 °C = 298 K.

**Are the Answers Reasonable?** For numerical calculations, ask yourself, "Is the answer too large or too small?" Judging the answers to such questions serves as a check on the arithmetic as well as on the method of obtaining the answer. For part **(a)**, we know that a Fahrenheit degree is about half the size of a Celsius degree, so 25 Celsius degrees should be about 50 Fahrenheit degrees, then we add 32 °F, so the Fahrenheit temperature should be approximately 32 °F + 50 °F = 82 °F. The answer of 77 °F is quite close.

For part **(b)**, we recall that 0 °C = 273 K. A temperature above 0 °C must be higher than 273 K. Our calculation, therefore, appears to be correct.

**PRACTICE EXERCISE 1.7**

What is 355 °C in °F? (*Hint:* What tool relates these two temperature scales?)

**PRACTICE EXERCISE 1.8**

Convert 55 °F to its Celsius temperature. What Kelvin temperature corresponds to 68 °F (expressed to the nearest whole kelvin unit)?

## 1.5 Making Measurements in the Laboratory

Chemistry is a science of measurements. Chemists measure masses of chemicals, volumes of solutions, and temperatures, pressures, and volumes of gases among many other things. This leads to the question, "Can we make the 'perfect measurement'"? The simplest measurement to make is to count a small group of objects. For instance, you can count the coins in your pocket or the number of cookies left in a package. The number of coins and cookies will be fairly small and you can count them without making a mistake. There will be no doubt about the result and we can call that a "perfect measurement." But, if there is a much larger number of cookies, perhaps more than 5000 in a barrel, the odds of your counting them correctly become vanishingly small and counting them one by one is a waste of time. In fact, statisticians have demonstrated that your result will be off by a number close to the square root of 5000, or ± 71. So much for making the perfect measurement.

### Determinate Errors

Let's consider what happens with a real measurement. To make a measurement, we need a tool. It can be as simple as a meter stick or as complex as the mass spectrometer shown in the On the Cutting Edge 0.2. Importantly, the tool must be used and maintained properly or the measurement will be incorrect. "Using the tool properly" means that there is a method for its use that the experimenter follows exactly. Any deviation from that method may result in incorrect results. Similarly, if the tool is not maintained properly, the results will not be correct. When this happens, we have a **determinate error** in the measurement. However, all the things that cause a determinate error can be corrected.

Determinate error can be measured and minimized by taking measurements with the tool on a standard sample with known properties. If you have a determinate error, the results of repeated measurements will be consistently either too high or too low compared to the standard. Also, the error (i.e., the difference between the measured value and the true value of the standard) will repeatedly have a positive or negative sign. To minimize the determinate error, measurements are made again after making small changes to the instrument or method to see if the results are closer to the correct value. When the measurement and the known value agree, the determinate error has been minimized.

**NOTE** Error = Measured Value – True Value

Error always has a sign, + or –, indicating that the results is either too high or too low.

## Uncertainty

Once the determinate error has been minimized, will our results will be perfect? Unfortunately, no. If repeated measurements are made, the measurements are not exactly the same. This is because measurements are made by using the calibration marks on the tool and then estimating the last digit between the calibration marks. The result is that the last (estimated) digit varies from person-to-person and from reading-to-reading. These differences are called **uncertainty**. Consider, for example, reading the same temperature from each of the two thermometers in **Figure 1.22**.

The marks on the left thermometer are one degree apart, and we can see that the temperature lies between 24 °C and 25 °C. When reading any scale, the last digit is estimated and recorded to the nearest tenth of the smallest scale division. Looking closely, therefore, we might estimate that the fluid column falls about 3/10 of the way between the marks for 24 and 25 degrees, so we can report the temperature to be 24.3 °C. However, it would be foolish to say that the temperature is exactly 24.3 °C. The last digit is only an estimate, and this thermometer might be read as 24.2 °C by one observer or 24.4 °C by another. Because of this, there is an uncertainty of ±0.1 °C in the measured temperature. We can express this by writing the temperature as 24.3 ± 0.1 °C.

The thermometer on the right has marks that are 1/10 of a degree apart, which allows us to estimate the temperature as 24.32 °C. In this case, we are estimating the hundredths place and the uncertainty is ±0.01 °C; therefore, the temperature is reported as 24.32 ± 0.01 °C. Notice that because the thermometer on the right is more finely graduated, we are able to obtain measurements with smaller uncertainties. We would have more confidence in temperatures read from the thermometer on the right in Figure 1.22 because it has more digits and a smaller amount of uncertainty. The reliability of a measurement is indicated by the number of digits used to represent it.

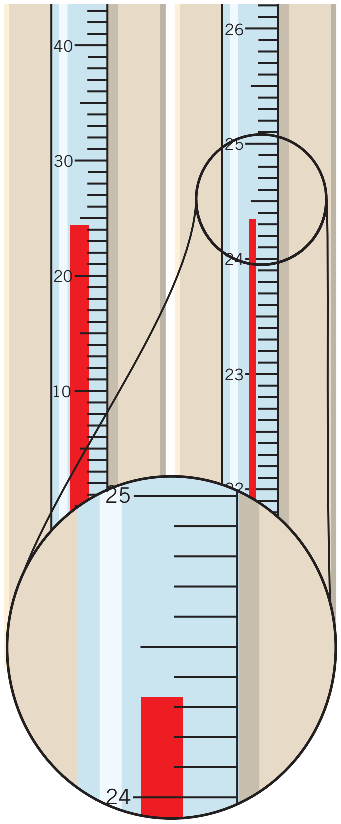



**FIGURE 1.22 Thermometers with different scales give readings with different precision.** The thermometer on the left has marks that are one degree apart, allowing the temperature to be estimated to the nearest tenth of a degree. The thermometer on the right has marks every 0.1 °C. This scale permits estimation of the hundredths place.

By convention in science, all digits in a measurement up to and including the first estimated digit are recorded. If a reading measured with the thermometer on the right seemed exactly on the 24 °C mark, we would record the temperature as 24.00 °C, not 24 °C, to show that the thermometer can be read to the nearest 1/100 of a degree.

Modern laboratory instruments such as balances and meter often have digital displays. If you read the mass of a beaker on a digital balance as 65.23 grams, everyone else observing the same beaker will see the same display and obtain exactly the same reading. There seems to be no uncertainty, or estimation, in reading this type of scale. However, scientists agree that the uncertainty is ±1/2 of the last readable digit because the actual mass of the beaker could range from 65.225 grams to 65.234 grams. Using this definition, our digital reading may be written as 65.230 ± 0.005 grams.[6]

[6]Some instructors may wish to maintain a uniform policy of assigning an uncertainty of ±1 in the last readable digit for both analog and digital scale readings.

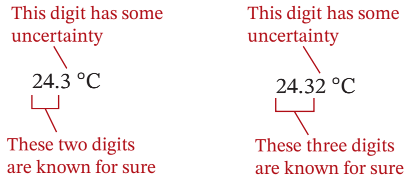
## Significant Figures

The concepts discussed above are so important that we have special terminology to describe numbers that come from measurements.

**TOOLS**
**Counting significant figures**

*Digits that result from a measurement such that only the digit farthest to the right is not known with certainty are called* **significant figures** (*or* **significant digits**).

The number of significant figures in a measurement is equal to the number of digits known for sure *plus* one that is estimated. Let's look at our two temperature measurements:




The first measurement, 24.3 °C, has three significant figures; the second, 24.32 °C, has four significant figures.

### When Are Zeros Significant?

Usually, it is simple to determine the number of significant figures in a measurement; we just count the digits. Thus, the number 3.25 has three significant figures and 56.215 has five of them. When zeros are part of a number, they can sometimes cause confusion.

We will use the following rules when zeros are involved:

*Zeros to the left of the first nonzero digit, called leading zeros, are never counted as significant.* For instance, a length of 2.3 mm is the same as 0.0023 m. Since we are dealing with the same measured value, its number of significant figures cannot change when we change the units. Both quantities have two significant figures.

*Zeros imbedded inside a number are always significant.* Thus, 2047 m and 3.096 g have four significant figures each because the zeros are imbedded.

*Trailing zeros are (a) always counted as significant if the number has a decimal point and (b) not counted as significant if the number does not have a decimal point.* Thus, 4.500 m and 630.0 g have four significant figures each because the zeros would not be written unless those digits were known to be zeros. The numbers 3200 and 20 have two and one significant figures, respectively.

There are times when a measurement or calculation requires the trailing zeros to be significant. The best way to handle this is to use **scientific notation** to write the number. For example, suppose you were told that a football game was attended by 45,000 people. If we want to report the number of sports fans as 45,000 give or take a thousand, we can write the rough estimate as $4.5 \times 10^4$. The 4.5 shows the number of significant figures and the $10^4$ tells us the location of the decimal point. On the other hand, suppose the sports fans were carefully counted using an aerial photograph, so that the count could be reported to be 45,000—give or take about 100 people. In this case, the value represents 45,000 $\pm$ 100 attendees and contains three significant figures, with the uncertain digit being the zero in the hundreds place. In scientific notation, this can be expressed as $4.50 \times 10^4$. This time the 4.50 shows three significant figures and an uncertainty of $\pm 0.01 \times 10^4$ or $\pm 100$ people. Thus, a simple statement such as "there were 45,000 people attending the game" is ambiguous. We can't tell how many significant digits the number has from the number alone. The best we can do is to say "45,000 has *at least* two significant figures."

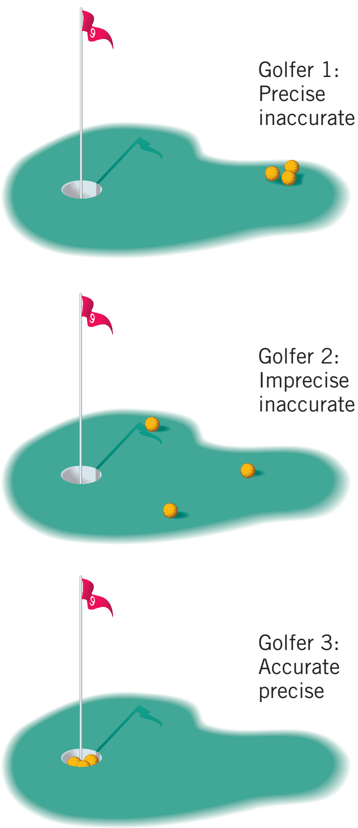
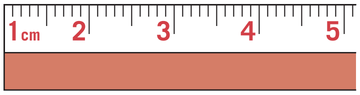



**FIGURE 1.23 The difference between precision and accuracy in the game of golf.** Golfer 1 hits shots that are precise (because they are tightly grouped), but the accuracy is poor because the balls are not near the target (the "true" value). Golfer 2 needs help. The shots are neither precise nor accurate. Golfer 3 wins the prize with shots that are precise (tightly grouped) and accurate (in the hole).

## Accuracy and Precision

Two words often used in reference to measurements are *accuracy* and *precision*. **Accuracy** *refers to how close a measurement is to the true or the accepted true value.* **Precision** *refers to how close repeated measurements come to their average*. Notice that the two terms are not synonyms, because accuracy is related to determinate errors and precision is related to random errors. A practical example of how accuracy and precision differ is illustrated in **Figure 1.23**.




How accurate would measurements be with this ruler?

**FIGURE 1.24 An improperly marked ruler.** This improperly marked ruler will yield measurements that are each wrong by one whole unit. The measurements might be precise, but the accuracy would be very poor.

For measurements to be accurate, the measuring device must be carefully calibrated (adjusted) with a standard reference so it gives correct readings. For example, to calibrate an electronic balance, a known reference mass is placed on the balance and a calibration routine within the balance is initiated. Once calibrated, the balance will give accurate readings, the accuracy of which is determined by the quality of the standard mass used. Standard reference masses, often "traceable" to the standard kilogram, can be purchased from scientific supply companies.

Precision refers to how closely repeated measurements of the same quantity come to each other. In general, the smaller the uncertainty (i.e., the "plus or minus" part of the measurement), the more precise the measurement. This translates as: *The more significant figures in a measured quantity, the more precise the measurement*.

We usually assume that a very precise measurement is also of high accuracy. We can be wrong, however, if our instruments are improperly calibrated. For example, the improperly marked ruler in **Figure 1.24** might yield measurements that vary by a hundredth of a centimeter ($\pm$0.01 cm), but all the measurements would be too large by 1 cm—a case of good precision but poor accuracy, and if this defect in the ruler was detected, it could be corrected.

## Significant Figures in Calculations

When several measurements are obtained in an experiment, they are usually combined in some way to calculate a desired quantity. For example, to determine the area of a rectangular carpet, we require two measurements, length and width, which are then multiplied to give the answer we want. To get some idea of how precise the area really is, we need a way to take into account the precision of the various values used in the calculation. To make sure this happens, we follow certain rules according to the kinds of arithmetic being performed.

### Multiplication and Division

*For multiplication and division, the number of significant figures in the answer should be equal to the number of significant figures in the least precise measurement.* The least precise measurement is the number with the fewest significant figures. Let's look at a typical problem involving some measured quantities.

**TOOLS**
**Significant figures: multiplication and division**

3 sig. figures 4 sig. figures

$$\frac{3.14 \times 2.751}{0.64} = 13$$

2 sig. figures

The result displayed on a calculator[7] is 13.49709375. However, the least precise factor, 0.64, has only two significant figures, so the answer should have only two. The correct answer, 13, is obtained by rounding-off the calculator answer. When we wish to round off a number at a certain point, we simply drop the digits that follow if the first of them is less than 5. Thus, 8.1634 rounds to 8.16 if we wish to have only two decimal places. If the first digit after the point of round-off is larger than 5, or if it is 5 followed by other nonzero digits, then we add 1 to the preceding digit. Thus 8.167 and 8.1653 both round to 8.17. Finally, when the digit after

[7]Calculators usually give too many significant figures. An exception is when the answer has zeros at the right that are significant figures. For example, an answer of 1.200 is displayed on most calculators as 1.2. If the zeros belong in the answer, be sure to write them down.

the point of roundoff is 5 and no nonzero digits follow the 5, then we drop the 5 if the preceding digit is even and add 1 if it is odd. Thus, 8.165 rounds to 8.16 and 8.17500 rounds to 8.18. When we multiply and divide measurements, the units of those measurements are multiplied and divided in the same way as the numbers.

**TOOLS**
**Significant figures: addition and subtraction**

### Addition and Subtraction

*For addition and subtraction, the answer should have the same number of decimal places as the quantity with the fewest number of decimal places.* As an example, consider the following addition of measured masses.

$$\begin{array}{r l} 3.247\text{ g} & \\ 41.36\text{ g} & \\ \underline{+125.2\text{ g}} & \longleftarrow\text{(This number has only 1 decimal place.)} \\ 169.8\text{ g} & \longleftarrow\text{(This answer has been rounded to 1 decimal place.)} \end{array}$$

In this calculation, the digits beneath the 6 and the 7 are unknown; they could be anything. (They're not necessarily zeros because if we *knew* they were zeros, then zeros would have been written there.) Adding an unknown digit to the 6 or 7 will give an answer that's also unknown, so for this sum we are not justified in writing digits in the second and third places after the decimal point. Therefore, we round the answer to the nearest tenth of a gram. Remember, we can only add and subtract numbers that have identical units, and the answer will have the same units in this case, grams. When a calculation has both addition (or subtraction) and multiplication (or division), do the addition/subtraction first to the correct number of significant figures and then perform the multiplication and division operations to the correct number of significant figures.

**TOOLS**
**Significant figures: exact numbers**

### Exact Numbers

Numbers that come from definitions, such as 12 in. = 1 ft, and those that come from a direct count, such as the number of people in a small room, have no uncertainty, and we can assume that they have an infinite number of significant figures. Therefore, exact numbers do not affect the number of significant figures in multiplication or division calculations.

#### PRACTICE EXERCISE 1.9

Perform the following calculations involving measurements and round the results so they have the correct number of significant figures and proper units. (*Hint:* Apply the rules for significant figures described in this section, and keep in mind that units behave as numbers do in calculations.)

**(a)** 21.0233 g + 21.0 g

**(b)** 10.0324 g ÷ 11.7 mL

**(c)** $\dfrac{14.25\text{ cm} \times 12.334\text{ cm}}{(2.223\text{ cm} - 1.04\text{ cm})}$

#### PRACTICE EXERCISE 1.10

Perform the following calculations involving measurements and round the results so that they are written to the correct number of significant figures and have the correct units.

**(a)** 54.183 g − 0.0278 g

**(b)** 10.0 g + 1.03 g + 0.243 g

**(c)** $43.4\text{ in.} \times \dfrac{1\text{ ft}}{12\text{ in.}}$

**(d)** $\dfrac{1.03\text{ m} \times 2.074\text{ m} \times 2.9\text{ m}}{12.46\text{ m} + 4.778\text{ m}}$

**(e)** 32.6 mm × 117.5 mm

**(f)** 15.7 mL + 25 mL

# 1.6 Dimensional Analysis

Earlier we mentioned that for numerical problems, we often do not have a specific equation to solve; instead, all we need to do is convert one set of units to another. After analyzing the problem and assembling the necessary information to solve it, scientists usually use a technique commonly called **dimensional analysis**, in which we follow the units, to help them perform the correct arithmetic. As you will see, this method also helps in analyzing the problem and selecting the tools needed to solve it.

**NOTE** To construct a valid conversion factor, the relationship between the units must be true. For example, the statement 3 ft = 41 in. is false. Although you might make a conversion factor out of it, any answers you would calculate are sure to be wrong. Correct answers require correct relationships between units.

## Conversion Factors

In dimensional analysis we treat a numerical problem as one involving a conversion of units (the dimensions) from one kind to another. To do this we use one or more *conversion factors* to change the units of the given quantity to the units of the answer:

$$(\text{given quantity}) \times (\text{conversion factor}) = (\text{desired quantity})$$

A **conversion factor** *is a fraction formed from a* valid *equality or equivalence between units and is used to switch from one system of measurement and units to another.* To illustrate, suppose we want to express a person's height of 72.0 inches in centimeters. To do this we need the relationship between the inch and the centimeter. We can obtain this from Table 1.4:

$$2.54 \text{ cm} = 1 \text{ in.} \quad (\text{exactly}) \tag{1.4}$$

**NOTE** The relationship between the inch and the centimeter is exact, so that the numbers in 1 in. = 2.54 cm have an infinite number of significant figures.

If we divide both sides of this equation by 1 in., we obtain a conversion factor.

$$\frac{2.54 \text{ cm}}{1 \text{ in.}} = \frac{1 \cancel{\text{in.}}}{1 \cancel{\text{in.}}} = 1$$

Notice that we have canceled the units from both the numerator and denominator of the center fraction, leaving the first fraction equaling 1. As mentioned earlier, *units behave just as numbers do in mathematical operations*; this is a key part of dimensional analysis. Let's see what happens if we multiply 72.0 inches, the height that we mentioned, by this fraction.

$$72.0 \cancel{\text{in.}} \times \left(\frac{2.54 \text{ cm}}{1 \cancel{\text{in.}}}\right) = 183 \text{ cm}$$
$$(\text{given quantity}) \times (\text{conversion factor}) = (\text{desired quantity})$$

Because we have multiplied 72.0 in. by something that is equal to 1, we haven't changed the person's height. We have, however, changed the units. Notice that we have canceled the unit inches. The only unit left is centimeters, which is the unit we want for the answer. The result, therefore, is the person's height in centimeters.

One of the benefits of dimensional analysis is that it often lets you know when you have done the *wrong* arithmetic. From the relationship in Equation 1.4, we can actually construct two conversion factors:

$$\frac{2.54 \text{ cm}}{1 \text{ in.}} \quad \text{and} \quad \frac{1 \text{ in.}}{2.54 \text{ cm}}$$

We used the first one correctly, but what would have happened if we had used the second by mistake?

$$72.0 \text{ in.} \times \frac{1 \text{ in.}}{2.54 \text{ cm}} = 28.3 \text{ in.}^2/\text{cm}$$

In this case, none of the units cancel. We get units of $\text{in.}^2/\text{cm}$ because inches times inches is inches squared. Even though our calculator may be very good at arithmetic, we've got the wrong answer. *Dimensional analysis lets us know we have the wrong answer because the units are wrong!*

There is a general strategy for all problems that involve a conversion. We always start with a given piece of data along with its units. Then we write down the desired units that we want the answer to have. This gives us the start and end points of our calculation. Then, all we need to do is find the correct conversion factors that will lead us from one set of units to the next. Sometimes this will require one conversion factor; other times it may take two or more to complete the conversion. The next examples illustrate this process.

## EXAMPLE 1.3 | Dimensional Analysis Applied to Metric Prefixes

The world's longest documented hair in 2004 belonged to Xie Qiuping of China and was measured at 5.627 m. She stopped cutting her hair in 1973, and over the course of 31 years, it grew to this length. Convert 5.627 m to millimeters (mm).

**Analysis:** We are asked to convert a number with meter units to another number that has millimeter units, a conversion of metric prefixes. This involves converting meters to millimeters (mm).

**Assembling the Tools:** To solve this problem, our tool will be a conversion factor that relates the unit meter to the unit millimeter. From Table 1.5, the table of decimal multipliers, the prefix "milli" means "$\times 10^{-3}$," so we can write

$$1 \text{ mm} = 10^{-3} \text{ m}$$

Notice that this relationship connects the units given to the units desired.

From this relationship, we can form two conversion factors.

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

We now have all the information we need to solve the problem.

**Solution:** We start with the given quantity and its units on the left and the units of the desired answer on the right.

$$5.627 \text{ m} = ? \text{ mm}$$

We have to cancel the unit meter, so we need to multiply by the conversion factor with meter in the denominator. Therefore, we select the one on the left as our tool. This gives

$$5.627 \cancel{\text{m}} \times \frac{1 \text{ mm}}{10^{-3} \cancel{\text{m}}} = 5.627 \times 10^3 \text{ mm or } 5627 \text{ mm}$$

Notice that we have expressed the answer to four significant figures because that is how many there are in the given quantity, 5.627 m. The equality that relates meters and millimeters involves exact numbers because it is a definition.

**Is the Answer Reasonable?** We know that millimeters are much smaller than meters, so 5.627 m must represent a lot of millimeters. We also know that conversions between prefixes will not change the 5.627 part of our number, just the power of 10. Our answer, therefore, makes sense.

## EXAMPLE 1.4 | Using Dimensional Analysis

A liter, which is slightly larger than a quart, is defined as 1 cubic decimeter (1 $dm^3$). How many liters are there in exactly 1 cubic meter (1 $m^3$)?

**Analysis:** Let's begin once again by stating the problem in equation form.

$$1\ m^3 = ?\ L$$

We don't have any direct conversions between cubic meters and liters. There was, however, a definition of the liter in terms of base metric units that we can combine with metric prefixes to do the job.

**Assembling the Tools:** The relationship between liters and cubic decimeters was given in Equation 1.1,

$$1\ L = 1\ dm^3$$

From the table of decimal multipliers, we also know the relationship between decimeters and meters,

$$1\ dm = 0.1\ m$$

but we need a relationship between cubic units. Since units undergo the same kinds of operations numbers do, we simply cube each side of this equation (being careful to cube both the numbers and the units).

$$(1\ dm)^3 = (0.1\ m)^3$$
$$1\ dm^3 = 0.001\ m^3 \qquad \textbf{(1.5)}$$

Notice how Equations 1.5 and 1.1 provide a path from the given units to those we seek. In the pathway below, the equalities that will be used to create the conversion factors are shown above the arrows, but we still need to make the appropriate ratio using the equalities.

$$m^3 \xrightarrow{0.001\ m^3 = 1\ dm^3} dm^3 \xrightarrow{1\ dm^3 = 1\ L} L$$

We need to identify the conversion factors needed for each step. For the first step, we will use Equation 1.5.

For the first step, from Equation 1.5, we can make two conversion factors

$$\frac{1\ dm^3}{0.001\ m^3} \quad \text{or} \quad \frac{0.001\ m^3}{1\ dm^3}$$

The conversion factor we want will be the one that cancels the $m^3$ units; that is the first one given. Recall that when units are on opposite sides of the fraction line, the units will cancel.

For the second step, from Equation 1.1, another two conversion factors can be made:

$$\frac{1\ L}{1\ dm^3} \quad \text{or} \quad \frac{1\ dm^3}{1\ L}$$

This time, we want to cancel $dm^3$ units; therefore, we choose the ratio where the $dm^3$ is in the opposite part of the fraction from where it starts. In this case, the $dm^3$ is in the numerator, so we choose the ratio with the $dm^3$ in the denominator.

Now we are ready to solve the problem.

**Solution:** The first step is to eliminate the units $m^3$, using the conversion factor we obtained from Equation 1.5.

$$1\ \cancel{m^3} \times \frac{1\ dm^3}{0.001\ \cancel{m^3}} = 1000\ dm^3$$

Then, using the conversion factor we developed from Equation 1.1, this will take us from $dm^3$ to L.

$$1000\ \cancel{dm^3} \times \frac{1\ L}{1\ \cancel{dm^3}} = 1000\ L$$

Thus, 1 $m^3$ = 1000 L.

When a problem involves the use of two or more conversion factors, often they can be "strung together" in a "chain calculation." For example, this problem can be set up as follows.

$$1\ \cancel{m^3} \times \frac{1\ \cancel{dm^3}}{0.001\ \cancel{m^3}} \times \frac{1\ L}{1\ \cancel{dm^3}} = 1000\ L$$

Since all of our conversion factors involve exact numbers and we were given exactly one cubic meter, our answer is an exact number also.

**Is the Answer Reasonable?** One liter is about a quart. A cubic meter is about a cubic yard. Therefore, we expect a large number of liters in a cubic meter, so our answer seems reasonable. (Notice here that in our analysis we have approximated the quantities in the calculation in units of quarts and cubic yards, which may be more familiar than liters and $m^3$. We get a feel for the approximate magnitude of the answer using our familiar units and then relate this to the actual units of the problem.)

## EXAMPLE 1.5 | Applying Dimensional Analysis to Non-SI Units

Some mountain climbers are susceptible to high-altitude pulmonary edema (HAPE), a life-threatening condition that causes fluid retention in the lungs. It can develop when a person climbs rapidly to heights greater than $2.5 \times 10^3$ meters (2500 m). What is this distance expressed in feet?

**Analysis:** The problem can be stated as

$$2.5 \times 10^3\ m = ?\ ft$$

in which we convert meters to feet. The route we take from meters to feet depends on the conversion factors we have available. Since Table 1.4 does not have a relationship between feet and meters we need to develop a sequence of conversions by looking at the units we can convert. One sequence may be:

$$2.5 \times 10^3\ m \xrightarrow{1\ m = 100\ cm} \text{centimeters} \xrightarrow{2.54\ cm = 1\ in.} \text{inches} \xrightarrow{12\ in. = 1\ ft} \text{feet}$$

or if we consult the inside back cover of this book, we find the conversion factor 1 yard = 0.9144 meters. Another multistep sequence could be

$$2.5 \times 10^3\ m \xrightarrow{0.9144\ m = 1\ yd} \text{yards} \xrightarrow{1\ yd = 3\ ft} \text{feet}$$

We will choose the first sequence for this example.

**Assembling the Tools:** The tools we will need can be found by looking in the appropriate tables as shown.

$$1\ cm = 10^{-2}\ m \qquad \text{(from Table 1.5)}$$
$$1\ in. = 2.54\ cm \qquad \text{(from Table 1.4)}$$
$$1\ ft = 12\ in.$$

**Solution:** Now we apply dimensional analysis by following our planned sequence of eliminating unwanted units to bring us to the units of the answer.

$$2.5 \times 10^3\ \cancel{m} \times \frac{1\ \cancel{cm}}{10^{-2}\ \cancel{m}} \times \frac{1\ \cancel{in.}}{2.54\ \cancel{cm}} \times \frac{1\ ft}{12\ \cancel{in.}} = 8.2 \times 10^3\ ft = 8200\ ft$$

This time the answer has been rounded to two significant figures because that's how many there were in the measured distance. Notice that the numbers 12 and 2.54 do not affect the number of significant figures in the answer because they are exact numbers derived from definitions.

As suggested in the analysis, this is not the only way we could have solved this problem. Other sets of conversion factors could have been chosen. For example, we could have followed the second sequence of conversions and used: 1 yd = 0.9144 m and 3 ft = 1 yd. Then the problem would have been set up as follows:

$$2500\ \cancel{m} \times \frac{1\ \cancel{yd}}{0.9144\ \cancel{m}} \times \frac{3\ ft}{1\ \cancel{yd}} = 8200\ ft$$

(Remember, this number has two significant figures.)

Many problems that you meet, just like this one, have more than one path to the answer. *The important thing is for you to be able to reason your way through a problem and find a set of relationships that can take you from the given information to the answer.* Dimensional analysis can help you search for these relationships if you keep in mind the units that must be eliminated by cancellation.

**Is the Answer Reasonable?** Let's do some approximate arithmetic to get a feel for the size of the answer. A meter is slightly longer than a yard, so let's approximate the given distance, 2500 m, as 2500 yd. In 2500 yd, there are $3 \times 2500 = 7500$ ft. Since the meter is a bit longer than a yard, our answer should be a bit longer than 7500 ft, so the answer of 8200 ft seems to be reasonable.

**PRACTICE EXERCISE 1.11**

A tiny house is one that is between 60 square feet to 400 square feet. A house that is 124 square feet would be considered a tiny house. Use dimensional analysis to convert an area of 124 square feet to square meters. (*Hint:* What relationships would be required to convert feet to meters?)

**PRACTICE EXERCISE 1.12**

Use dimensional analysis to perform the following conversions: **(a)** 3.00 $yd^3$ to $in.^3$, **(b)** 1.25 km to cm, **(c)** 3.27 oz to g, **(d)** 20.2 mi/gal to km/L.

## Equivalencies

Up to now we've constructed conversion factors from relationships that are literally equalities. We can also make conversion factors from expressions that show how one thing is *equivalent* to another. For instance, if you buy a pair of shoes for \$85, we can say you converted \$85 into a pair of shoes or \$85 is equivalent to a pair of shoes. We would write this as

$$\$85 \Leftrightarrow 1 \text{ pair of shoes}$$

**NOTE** The equivalence symbol, ⇔, acts just like an equal sign.

where the symbol ⇔ is read as "is equivalent to." Mathematically, this equivalence sign works the same as an equal sign and we can construct two conversion factors:

$$\frac{\$85}{1 \text{ pair of shoes}} \quad \text{or} \quad \frac{1 \text{ pair of shoes}}{\$85}$$

that allow us to convert from shoes to dollars or from dollars to shoes. We can use the first conversion factor to determine how much it would cost to buy three pairs of shoes:

$$3 \cancel{\text{ pairs of shoes}} \left( \frac{\$85}{1 \cancel{\text{ pair of shoes}}} \right) = \$255$$

The "units" pairs of shoes cancel, leaving just the cost of the shoes, \$255.

# 1.7 Density and Specific Gravity

In our earlier discussion of properties, we noted that intensive properties are useful for identifying substances. One of the interesting things about extensive properties is that if you take the ratio of two of them, the resulting quantity is usually independent of sample size. In effect, the sample size cancels out and the calculated quantity becomes an intensive property. One useful property obtained this way is **density**, *which is defined as the ratio of an object's mass to its volume*. Using the symbols $d$ for density, $m$ for mass, and $V$ for volume, we can express this mathematically as

**TOOLS** Density

$$d = \frac{m}{V} \tag{1.6}$$

Notice that to determine an object's density, we make two measurements, mass and volume.

## EXAMPLE 1.6 | Calculating Density

A sample of blood completely fills an 8.20 $cm^3$ vial. The empty vial has a mass of 10.30 g and the filled vial has a mass of 18.91 g. What is the density of blood in units of $g/cm^3$?

**Analysis:** In this problem, the mass and volume of blood are related through its density. We are given the volume of the blood, and its mass can be calculated from the mass of the full vial and the mass of the empty vial. Once we have the mass of blood and its volume, we can use the definition of density to solve the problem.

**Assembling the Tools:** The law of conservation of mass is one tool we need. It can be stated as follows:

$$\text{mass of full vial} = \text{mass of empty vial} + \text{mass of blood}$$

The other tool needed is the definition of density, given by Equation 1.6,

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

**Solution:** The mass of the blood is the difference between the masses of the full and empty vials:

$$\text{mass of blood} = \text{mass of full vial} - \text{mass of empty vial}$$
$$\text{mass of blood} = 18.91\text{ g} - 10.30\text{ g} = 8.61\text{ g}$$

To determine the density, we simply take the ratio of mass to volume:

$$\text{density} = \frac{m}{V} = \frac{8.61\text{ g}}{8.20\text{ cm}^3} = 1.05\text{ g cm}^{-3}\text{ or }1.05\text{ g/cm}^3$$

This could also be written as density $= 1.05$ g/mL, because $1\text{ cm}^3 = 1$ mL.

**Is the Answer Reasonable?** First, the answer has the correct units, so that's encouraging. In the calculation we are dividing 8.61 by a number that is slightly smaller, 8.20. The answer should be slightly larger than one, which it is, so a density of 1.05 $g/cm^3$ seems reasonable.

### PRACTICE EXERCISE 1.13

Polystyrene foam is an excellent insulator, is lightweight, and does not absorb moisture, making it excellent material for insulated foam cups. A 15.0 mL sample of polystyrene used in insulated foam cups weighs 0.244 g. What is the density of this polystyrene? (*Hint:* What is the relationship between mass and volume in density calculations?)

### PRACTICE EXERCISE 1.14

A crystal of salt was grown and found to have a mass of 0.547 g. When the crystal was placed in a graduated cylinder that contained 5.70 mL oil, the volume increased to 5.95 mL. What is the density of salt? Give your answer in $g\ cm^{-3}$ since the density of solids are reported in $g\ cm^{-3}$.

**NOTE** In chemistry, reference data are commonly tabulated at 25 °C, which is close to room temperature. Biologists often carry out their experiments at 37 °C because that is our normal body temperature.

**NOTE** Although the density of water varies slightly with temperature, it is very close to 1.00 $g/cm^3$ when the temperature is close to room temperature, 25 °C.

Each pure substance has its own characteristic density (**Table 1.6**). Gold, for instance, is much more dense than iron. Each $cm^3$ of gold has a mass of 19.3 g, so its density is 19.3 $g/cm^3$, while iron has a density of 7.86 $g/cm^3$. By comparison, the density of water is 1.00 $g/cm^3$, and the density of air at room temperature is about 0.0012 $g/cm^3$.

Most substances, such as the fluid in the bulb of a thermometer, expand slightly when they are heated, so the amount of matter packed into each $cm^3$ is less. Therefore, density usually decreases slightly with increasing temperature.[8] For solids and liquids, the size of this change is small, as you can see from the data for water in **Table 1.7**. When only two or three significant figures are required, we can often ignore the variation of density with small changes in temperature.

[8]Liquid water behaves oddly. Its maximum density is at 4 °C, so when water at 0 °C is warmed, its density increases until the temperature reaches 4 °C. As the temperature is increased further, the density of water gradually decreases.

**TABLE 1.6 Densities of Some Common Substances at Room Temperature[a]**

| Substance | Density ($g/cm^3$) |
|---|---|
| Water | 1.00 |
| Aluminum | 2.70 |
| Iron | 7.87 |
| Silver | 10.5 |
| Gold | 19.3 |
| Glass[b] | 2.4 |
| Air[c] | 0.0012 |

[a]P. J. Linstrom and W. G. Mallard, Eds., *NIST Chemistry WebBook, NIST Standard Reference Database Number 69*, National Institute of Standards and Technology, Gaithersburg MD, 20899, http://webbook.nist.gov (retrieved June, 2021).

[b]Plate glass

[c]Dry, near sea level

**TABLE 1.7 Density of Water as a Function of Temperature[a]**

| Temperature (°C) | Density ($g/cm^3$) |
|---|---|
| 10 | 0.9997025 |
| 15 | 0.9991026 |
| 20 | 0.9982072 |
| 25 | 0.9970476 |
| 30 | 0.9956495 |
| 50 | 0.9880350 |
| 100 | 0.9583675 |

[a]P. J. Linstrom and W. G. Mallard, Eds., *NIST Chemistry WebBook, NIST Standard Reference Database Number 69*, National Institute of Standards and Technology, Gaithersburg MD, 20899, http://webbook.nist.gov (retrieved June, 2021).

## Density as a Conversion Factor

A useful property of density is that it provides a way to convert between the mass and volume of a substance. It defines the relationship, or equivalence, between the substance's mass and its volume. For instance, the density of gold (19.3 $g/cm^3$) tells us that 19.3 g of the metal is equivalent to a volume of 1.00 $cm^3$,

$$19.3 \text{ g gold} \Leftrightarrow 1.00 \text{ cm}^3 \text{ gold}$$

In setting up calculations for dimensional analysis, this equivalence can be used to form the two conversion factors:

$$\frac{19.3 \text{ g gold}}{1.00 \text{ cm}^3 \text{ gold}} \quad \text{and} \quad \frac{1.00 \text{ cm}^3 \text{ gold}}{19.3 \text{ g gold}}$$

The following example illustrates how we use density in calculations.

### EXAMPLE 1.7 | Calculations Using Density

Seawater has a density of about 1.03 g/mL. **(a)** What mass of seawater would fill a sampling vessel to a volume of 225 mL? **(b)** What is the volume, in mL of 45.0 g of seawater?

**Analysis:** For both parts of this problem, we are relating the mass of a material to its volume. Density provides a direct relationship between mass and volume, so this should be a one-step conversion.

**Assembling the Tools:** Density, Equation 1.6, is the only tool that we need to convert between these two quantities. We write the equivalence for this problem as

$$1.03 \text{ g seawater} \Leftrightarrow 1.00 \text{ mL seawater}$$

From this relationship we can construct two conversion factors. These will be the tools we use to obtain the answers.

$$\frac{1.03 \text{ g seawater}}{1.00 \text{ mL seawater}} \quad \text{and} \quad \frac{1.00 \text{ mL seawater}}{1.03 \text{ g seawater}}$$

**Solution to (a):** The question can be restated as:

$$225 \text{ mL seawater} \Leftrightarrow ? \text{ g seawater}$$

We need to eliminate the unit *mL seawater*, so we choose the conversion factor on the left as our tool.

$$225 \cancel{\text{mL seawater}} \times \frac{1.03 \text{ g seawater}}{1.00 \cancel{\text{mL seawater}}} = 232 \text{ g seawater}$$

Thus, 225 mL of seawater has a mass of 232 g.

**Solution to (b):** The question is:

$$45.0 \text{ g seawater} \Leftrightarrow ? \text{ mL seawater}$$

This time we need to eliminate the unit *g seawater*, so we use the conversion factor on the right as our tool.

$$45.0 \cancel{\text{g seawater}} \times \frac{1.00 \text{ mL seawater}}{1.03 \cancel{\text{g seawater}}} = 43.7 \text{ mL seawater}$$

Thus, 45.0 g of seawater has a volume of 43.7 mL.

**Are the Answers Reasonable?** Notice that the density tells us that 1 mL of seawater has a mass of slightly more than 1 g. Thus, for part **(a)**, we might expect that 225 mL of seawater should have a mass slightly more than 225 g. Our answer, 232 g, is reasonable. For part **(b)**, 45 g of seawater should have a volume not too far from 45 mL, so our answer of 43.7 mL is the right size.

**PRACTICE EXERCISE 1.15**

A gold-colored metal object has a mass of 365 g and a volume of 22.12 $cm^3$. Is the object composed of pure gold? (*Hint:* How does the density of the object compare with that of pure gold?)

**PRACTICE EXERCISE 1.16**

A certain metal alloy has a density of 12.6 $g/cm^3$. How many pounds would 0.822 $ft^3$ of this alloy weigh? (*Hint:* What is the density of this alloy in units of $lb/ft^3$?)

## Specific Gravity

Density is the ratio of mass to volume of a substance. We used units of grams for the mass and cubic centimeters for the volume in our examples. This is entirely reasonable since most liquids and solids have densities between 0.5 and 20 $g/cm^3$. There are many different units for mass. Kilograms, grams, micrograms, pounds, and ounces come to mind, and there are undoubtedly more. Volume also has many possible units: $cm^3$, liters, ounces, gallons, and so on. From those listed, there are 20 different possible ratios of mass and volume units for density. If each set of units required a separate table, this would result in each substance having 20 different densities listed in tables.

The concept of *specific gravity solves this problem.* The **specific gravity** for a substance is simply the density of that substance divided by the density of water. The units for the two densities must be the same so that specific gravity will be a *dimensionless number*. In addition other experimental conditions, such as temperature, for determining the two densities must be the same.

**TOOLS**
**Specific gravity**

$$\text{Specific gravity} = \frac{\text{density of substance}}{\text{density of water}} \tag{1.7}$$

To use the specific gravity, the scientist simply selects the specific gravity of a substance and then multiplies it by the density of water that has the units desired. Now we can have a relatively compact table that lists the specific gravity for our chemical substances and then a second short table of the density of water, perhaps in the 20 different units suggested above.

## 1.3 Chemistry Outside the Classroom

### Density and Wine

Density, or specific gravity, is one of the basic measurements in the wine-making process. The essential chemical reaction is that yeast cells feed on the sugars in the grape juice, and one of the products is ethanol (ethyl alcohol) and the other product is carbon dioxide. This is the fermentation process. The two main sugars in grapes are glucose and fructose. The fermentation reaction for these sugars is

pixel2013/2420 images/Pixabay

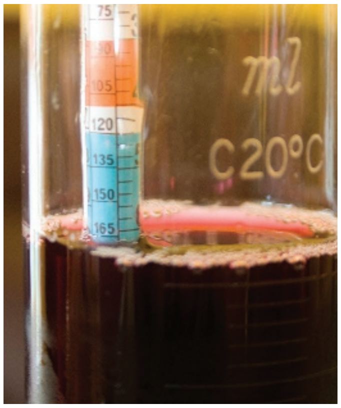
$$C_6H_{12}O_6(aq) \longrightarrow 2CH_3CH_2OH(aq) + 2CO_2(g)$$

If there is too little sugar, the amount of alcohol in the product will be low. Too much sugar will result in a maximum amount of alcohol, about 13%, but there will be leftover sugar. Wines that are sweet with leftover sugars or low in alcohol are not highly regarded.

To make high-quality wines, the natural sugar content of the grapes is monitored carefully, toward the end of the growing season. When the sugar content reaches the optimum level, the grapes are harvested, crushed, and pressed to start the fermentation process.

A very quick way to determine the sugar content of the grapes is to determine the density of the juice. The higher the density, the higher the sugar content. Vintners use a simple device called a hydrometer. The **hydrometer** is a weighted glass bulb with a graduated stem. Placed in a liquid and given a slight spin, the hydrometer will stay centered in the graduated cylinder and sink to a level that is proportional to the density. After the density is read from the hydrometer scale, the vintner consults a calibration chart to find the corresponding sugar content. Once the fermentation ends, the alcohol content must also be determined. Again, this is done by measuring the density. To do this, a sample of wine is distilled by heating it to boiling and condensing the alcohol and water that are vaporized. When the required amount of liquid has been condensed, its density is determined using another hydrometer. Using calibration tables, the density reading is converted into the percentage alcohol.

warren mcconnaughie/Alamy Stock Photo




**Hydrometer** used for measuring sugar content in crushed grapes before fermentation.


Paul Silverman/Fundamental Photographs

**Hydrometer** used for measuring alcohol content in a distilled wine sample after fermentation.

## EXAMPLE 1.8 | Using Specific Gravity

Concentrated sulfuric acid is sold in bottles with a label that states that the specific gravity at 25 °C is 1.84. How many cubic feet of sulfuric acid will weigh 55.5 pounds?

**Analysis:** We rewrite the question as an equation,

$$55.5 \text{ lb sulfuric acid} = ?\ \text{ft}^3 \text{ sulfuric acid}$$

We saw how to use density to convert mass to volume. The specific gravity needs to be converted to density units, preferably pounds and cubic feet.

**Assembling the Tools:** Using Equation 1.7 which defines specific gravity,

$$\text{specific gravity} = \frac{\text{density of substance}}{\text{density of water}}$$

will allow us to calculate the density of sulfuric acid. **Table 1.8** provides a list of densities.We will then use the density as a conversion factor as we did in Example 1.7.

**Solution:** Rearranging the specific gravity equation, we get

$$\text{density sulfuric acid} = (\text{specific gravity}) \times (\text{density of water}) = 1.84 \times 62.4 \text{ lb/ft}^3$$
$$\text{density sulfuric acid} = 114.8 \text{ lb/ft}^3$$

**TABLE 1.8**

| Density of Water Expressed with Different Units |
|---|
| 1.00 g/mL |
| 1000 kg/m$^3$ |
| 62.4 lb/ft$^3$ |
| 8.34 lb/gallon |

We now perform the conversion with the conversion factor from the density,

$$55.5 \cancel{\text{lb}} \text{ sulfuric acid} \times \frac{1 \text{ ft}^3}{114.8 \cancel{\text{lb}}} = 0.483 \text{ ft}^3$$

**Is the Answer Reasonable?** We can do some quick estimates. The density seems correct: since the specific gravity is approximately 2 and the density of water is approximately 60, an answer near 120 is expected. For the second part, we see that 114 is about twice the size of 55.5, so we expect an answer of about 0.5. Our answer is very close to that, suggesting that our answers are reasonable.

**PRACTICE EXERCISE 1.17**

Table wines have a minimum specific gravity of 1.090. If a barrel of wine contains 79 gallons, what is the mass of the wine in the barrel in pounds? (*Hint:* What is the relationship between specific gravity and density?)

**PRACTICE EXERCISE 1.18**

Diabetes can cause a decrease in the specific gravity of urine. If normal urine has a specific gravity between 1.002 and 1.030, and a sample of urine has a density of 1.008 oz/liq. oz., does the urine sample have a low specific gravity?

## Importance of Reliable Measurements

We saw earlier that substances can be identified by their properties. If we are to rely on properties such as density for identification of substances, it is very important that our measurements be reliable. We must have some idea of what the measurements' accuracy and precision are.

The importance of accuracy is obvious. If we have no confidence that our measured values are close to the true values, we certainly cannot trust any conclusions that are based on the data we have collected (**Table 1.9**).

**NOTE** The carat system for gold states that pure gold is referred to as 24 carat gold. Gold that is 50% gold by mass will be 12 carat. Gold that is less than 24 carats is usually alloyed with cheaper metals such as silver or copper.

Precision of measurements can be equally important. For example, suppose we have a gold ring and we want to determine its purity—that is, 24 carat gold, 22 carat gold, or 18 carat gold. We could determine the mass of the ring and then its volume, compute the density of the ring, and then compare our experimental density with the density of 24 carat gold (19.3 g/mL), 22 carat gold (17.7 g/mL), or 18 carat gold (17 g/mL). If we used a graduated cup measure and a kitchen scale, we would obtain a volume of 1.0 mL and a mass of 18 g, and the density of the ring would then be 18 g/mL. On the other hand, if we measured the mass of the ring with a laboratory balance capable of measurements to the nearest ±0.001 g and used volumetric glassware to measure the volume, we would find that the mass of the ring is 18.153 g and the volume is 1.03 mL. The density would be 17.6 g/mL, to the correct number of significant figures.

**TABLE 1.9 Uncertainty Depends on the Quality of the Measuring Equipment**

| | Graduated Cup/ Kitchen Scale | Laboratory Balance/ Volumetric Glassware |
|---|---|---|
| Mass (g) | 18 | 18.153 |
| Volume (mL) | 1.0 | 1.03 |
| Density (g/mL) | 18 | 17.6 |
| Number of significant figures | 2 | 3 |
| Error (g/mL) | ±1 | ±0.1 |
| Range of possible values (g/mL) | 17 – 19 | 17.5 – 17.7 |

If we compare this density to the densities of gold, rounded to two significant figures, 24 carat gold has a density of 19 g/mL and 18 carat gold has a density of 17 g/mL; then the ring could be anywhere from 18 to 24 carat gold. Instead, using the more precise measurements, the density of the ring is 17.6 g/mL, which matches the density of 22 carat gold, which has a density of 17.7 g/mL. This is one example of why precise and accurate measurements are important. Every minute of every day there are multi-million-dollar transactions made in the business world that depend on accurate and precise measurements such as this.

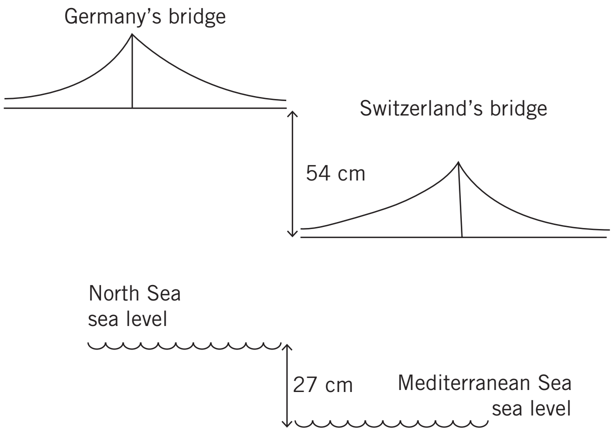
## Integrated Chemistry Example

What is sea level? The answer depends on which country responds. Germany measures sea level in relation to the North Sea, while Switzerland uses the Mediterranean Sea. In 2002, Germany and Switzerland were building a bridge in Laufenburg, a town that straddles Germany and Switzerland. The builders knew that there was a 27 cm difference between the two versions of sea level; however, when the two parts of the bridge met, there was a 54 cm difference between the sides. In the end, the Germans lowered its side of the bridge by the 54 cm. Explain how could this have happened, and what could be done in the future to avoid the situation?

**Analysis:** We are being asked to develop an explanation of how the mistake was made, and then come up with a way to keep this from happening again.

**Assembling the Tools:** This question might require a drawing to illustrate what happened and then a list of a few ideas which could be used to keep the problem from happening again.

**Solution:** For the first question, let's draw a sketch of the bridge, sea level, and their heights.




Germany's side of the bridge was 54 cm higher than Switzerland's side. The North Sea is 27 cm higher than the Mediterranean Sea. It looks as if Germany added the difference in sea levels to their calculations to make their side of the bridge higher.

$$27 \text{ cm} + 27 \text{ cm} = 54 \text{ cm}$$

If Germany had subtracted the 27 cm in their calculations, then the two sides would have met in alignment.

In the future, this type of mistake could be prevented by drawing a diagram of the situation and showing what would happen when the calculation is made. Another preventative measure would be to use the same sea level. Finally, instead of sea level, GPS measurements could be used.

**Is the Answer Reasonable:** Our diagram could be adapted to subtracting 27 cm two times from the German side to have the two bridge levels the same. Also, we have three different ways to prevent this issue in the future, and if all three are used, then any mistakes would be caught.

## Integrated Chemistry Problems

**A.** As you work through problems in this book, you will realize that many of the problems will require dimensional analysis. Draw a scheme that you can use as a model for solving dimensional analysis problems. In your scheme include notations for where the data comes from, what information you might have to look up, and how will you know that your answer is correct.

**B.** The SI units use decimal multipliers to form larger or smaller units. Most of the decimal multipliers change the units by a factor of 1000. Develop an explanation for why 1000 would be a reasonable number for the decimal multipliers.

**C.** When a person travels to another country from the United States, both the currency and the units of measure change. For example, when a person from Miami visits Rome, Italy, and wants to purchase fruit, say peaches, they need to use the Euro, €, and at the same time, the price is given per kilogram of peaches. Devise a way to calculate the cost of peaches in dollars/pound that can be done quickly and without the use of a calculator or looking up the value.

**D.** When the definition of the base units changed to be calculated from universal constants, this eliminated the need for a standard kilogram. In the 1980s, the standard meter was replaced by the definition of the meter, which is exactly the distance light travels in a vacuum in 1/299,792,458 of a second. The standard kilogram and the standard meter were actual objects that were kept in a protective environment and only used rare occasions to make secondary standards. Considering that these standards were physical objects, what would be the effect of temperature, pressure, or oxygen on these objects?

**E.** When a measurement is expressed, the units must always accompany the number. For example, Canada switched to the metric system in 1970, and, in 1983, an Air Canada plane had a malfunctioning fuel gauge. The crew used measuring "dipsticks" to check how much fuel the plane took on during refueling. They converted the volume measured to weight, and later, in flight, the plane ran out of fuel. The pilot was able to land the plane safely on the Gimli, Manitoba runway. Based on the units they used and the ones they could have used, explain why the plane ran out of fuel.

**F.** List the physical and chemical properties mentioned in this chapter. What additional physical and chemical properties can you think of to extend this list?

# Summary

## Organized by Learning Objective

### Explain the scientific method.

**Chemistry** is a science that studies the properties and composition of **matter**, which is defined as anything that has **mass** and occupies space. Chemistry employs the scientific method in which **observations** are used to collect **empirical facts**, or **data**, that can be summarized in **scientific laws**. **Models** of nature begin as **hypotheses** that mature into **theories** when they survive repeated testing.

### Classify matter.

An **element**, which is identified by its **chemical symbol**, cannot be decomposed into something simpler by a **chemical reaction**. Elements combine in fixed proportions to form **compounds**. Elements and compounds are **pure substances** that may be combined in varying proportions to give **mixtures**. If a mixture has two or more **phases**, it is **heterogeneous**. A one-phase **homogeneous** mixture is called a **solution**. Formation or separation of a mixture into its components can be accomplished by a **physical change**. Formation or decomposition of a compound takes place by a **chemical change** that changes the chemical makeup of the substances involved.

### Identify which physical and chemical properties are important to measure.

**Physical properties** are measured without changing the chemical composition of a sample. **Solid, liquid,** and **gas** are the most common **states of matter.** The properties of the states of matter can be related to the different ways the individual atomic-size particles are organized. A **chemical property** describes a chemical reaction a substance undergoes. **Intensive properties** are independent of sample size; **extensive properties** depend on sample size.

### Describe measurements and the use of SI units.

**Qualitative observations** lack numerical information, whereas **quantitative observations** require numerical measurements. The units used for scientific measurements are based on the set of seven **SI base units**, which can be combined to give various **derived units**. These all can be scaled to larger or smaller sized units by applying **decimal multipliers**. Convenient units for length and volume are, respectively, **centimeters** or **millimeters** and **liters** or **milliliters**. **Mass** is a measure of the amount of matter in an object and is measured with a **balance** and is expressed in units of **kilograms** or **grams**. Temperature is measured in units of **degrees Celsius**

(or **Fahrenheit**) using a thermometer. For many calculations, temperature must be expressed in **kelvins (K)**. The zero point on the **Kelvin temperature scale** is called **absolute zero**.

### Summarize how error and uncertainty in measurements occur and are reported using significant figures.

The **precision** of a measured quantity is expressed by the number of **significant figures** that it contains, which equals the number of digits known for sure plus the first one that possesses some uncertainty. Measured values are **precise** if they contain many significant figures and differ from each other by small amounts. A measurement is **accurate** if its value lies very close to the true value. When measurements are combined in calculations, rules help us determine the correct number of significant figures in the answer. **Exact numbers** are considered to have an infinite number of significant figures.

### Use dimensional analysis effectively.

**Dimensional analysis** is based on the ability of units to undergo the same mathematical operations as numbers. **Conversion factors** are constructed from *valid relationships* between units. These relationships can be either equalities or **equivalencies** (indicated by the symbol ⇔) between units.

### Determine the density of substances and use density as a conversion factor.

**Density** is an intensive property equal to the ratio of a sample's mass to its volume. Besides serving as a means for identifying substances, density provides a conversion factor that relates mass to volume. **Specific gravity** is the ratio of the density of a substance to the density of water and is a **dimensionless** quantity. This serves as a convenient way to have density data available in a large number of mass and volume units.

# Tools for Problem Solving

**TOOLS** The following tools were introduced in this chapter. Study them carefully so you can select the appropriate tool when needed.

**Base SI units** (Section 1.4) The eight base units of the SI system are used to derive the units for all scientific measurements.

**SI prefixes** (Section 1.4) The prefixes are used to create larger and smaller units.

**Units in laboratory measurements** (Section 1.4) Conversion for commonly used laboratory measurements.

**Length:** 1 m = 100 cm = 1000 mm
**Volume:** 1 L = 1000 mL = 1000 $cm^3$

**Temperature conversions: Celsius to Fahrenheit and Celsius to Kelvin** (Section 1.4)

$$t_F = \left(\frac{9\ °F}{5\ °C}\right)t_C + 32\ °F$$

$$T_K = (t_C + 273.15)\,°C\left(\frac{1\ K}{1\ °C}\right)$$

**Counting significant figures** (Section 1.5)

- All nonzero digits are significant.
- Zeros to the *left* of the first nonzero digit are never significant.
- Zeros between significant digits, imbedded zeros, are significant.
- Zeros on the end of a number (a) with a decimal point are significant; (b) those without a decimal point are assumed not to be significant. (To avoid confusion, scientific notation should be used.)

**Significant figures: multiplication and division** (Section 1.5) Round the answer to the same number of significant figures as the factor with the fewest significant figures.

**Significant figures: addition and subtraction** (Section 1.5) Round the answer to the same number of decimal places as the quantity with the fewest decimal places.

**Significant figures: exact numbers** (Section 1.5) Numbers from definitions, coefficients, and subscripts are exact and do not affect the number of significant figures of a calculation.

**Density** (Section 1.7)

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

**Specific gravity** (Section 1.7)

$$\text{Specific gravity} = \frac{d_{\text{substance}}}{d_{\text{water}}}$$

**Review Problems are presented in pairs. Answers to problems whose numbers appear in blue are given in Appendix B. More challenging problems are marked with an asterisk*. Additional questions and resources are available for assignment in WileyPLUS.**

# Review Questions

### Laws and Theories: The Scientific Method

**1.1** After some thought, give two reasons why a course in chemistry will benefit you in the pursuit of your particular major.

**1.2** Although there is no prescribed sequence of steps that scientists use in scientific investigations, describe the essence of what scientists do to discover new knowledge.

**1.3** What is the difference between **(a)** a law and a theory, **(b)** an observation and a conclusion, and **(c)** an observation and data?

**1.4** Can a theory be proved to be correct? Can a theory be proved to be wrong?

### Matter and Its Classifications

**1.5** What is matter? Which of the following are examples of matter? **(a)** air, **(b)** an idea, **(c)** a bowl of soup, **(d)** a squirrel, **(e)** sodium chloride, **(f)** the sound of an explosion

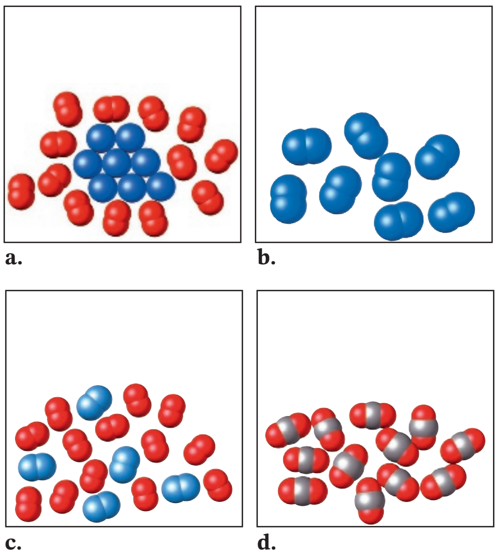
**1.6** Describe the following terms commonly used in chemistry. **(a)** element, **(b)** compound, **(c)** mixture, **(d)** homogeneous, **(e)** heterogeneous, **(f)** phase, **(g)** solution.

**1.7** Which kind of change, chemical or physical, is needed to change a compound into its elements?

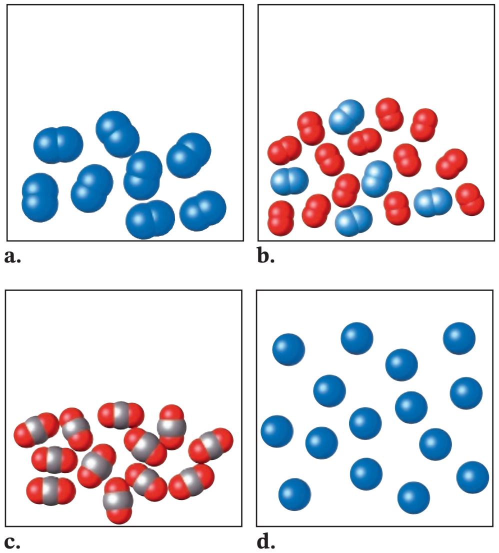
**1.8** What is the chemical symbol for each of the following elements? See how many symbols you can identify before looking at other sources. **(a)** fluorine, **(b)** selenium, **(c)** nickel, **(d)** argon, **(e)** lithium, **(f)** phosphorus, **(g)** iodine, **(h)** gallium, **(i)** mercury, **(j)** manganese

**1.9** What is the name of each of the following elements? See how many element names you can identify before looking at other sources. **(a)** Na, **(b)** Zn, **(c)** Si, **(d)** Sn, **(e)** Mg, **(f)** W, **(g)** Co, **(h)** Al, **(i)** O, **(j)** N

**1.10** For each of the following molecular pictures, state whether it represents a pure substance or a mixture. For pure substances, state whether it represents an element or a compound. For mixtures, state whether it is homogeneous or heterogeneous.




**1.11** Consider the following four samples of matter.




**(a)** Which sample(s) consist of only one element? **(b)** Which sample(s) consist of only one compound? **(c)** Which sample(s) consist of diatomic molecules?

### Physical and Chemical Properties

**1.12** Give two everyday examples of physical changes and two examples of chemical changes. What is distinctive about a physical change as compared to a chemical change.

**1.13** "A sample of calcium (an electrically conducting white metal that is shiny, relatively soft, melts at 850 °C, and boils at 1440 °C) was placed into liquid water that was at 25 °C. The calcium reacted slowly with the water to give bubbles of gaseous hydrogen and a solution of the substance calcium hydroxide." Based on this description, what physical changes and what chemical changes occurred?

**1.14** In places like Saudi Arabia, freshwater is scarce and is recovered from seawater. When seawater is boiled, the water evaporates and the steam can be condensed to give pure water that people can drink. If all the water is evaporated, solid salt is left behind. Are the changes described here chemical or physical?

**1.15** How does a chemical property differ from a physical property?

**1.16** Distinguish between an extensive and an intensive property.

**1.17** Determine whether each of the following is an intensive or extensive property, justify your reasoning, and give examples for each.

| | |
|---|---|
| **(a)** mass | **(e)** physical state |
| **(b)** boiling point | **(f)** density |
| **(c)** color | **(g)** volume |
| **(d)** surface area | |

**1.18** Describe one or more physical properties of each state of matter that distinguishes it from the other states of matter:

**(a)** gas **(b)** liquid **(c)** solid

**1.19** Density is an intensive property, as is temperature. Why are these intensive properties? Length and mass are extensive properties, why are these extensive?

**1.20** Besides density and temperature, what are two more intensive properties. Besides length and mass, what are two more extensive properties?

### Measurement of Physical and Chemical Properties

**1.21** Why must measurements always be written with a unit?

**1.22** What is the only SI base unit that includes a prefix?

**1.23** Which SI units are mainly used in chemistry?

**1.24** What are derived units? Give two examples of derived units.

**1.25** Answer this question, first, without consulting Table 1 and then complete the answers with the help of the table. What is the meaning of each of the following prefixes? What abbreviation is used for each prefix?

**(a)** centi
**(b)** milli
**(c)** kilo
**(d)** micro
**(e)** nano
**(f)** pico
**(g)** mega

**1.26** What reference points do we use in calibrating the scale of a thermometer? What temperature on the Celsius scale do we assign to each of these reference points?

**1.27** In each pair, which is larger:

**(a)** A Fahrenheit degree or a Celsius degree?
**(b)** A Celsius degree or a kelvin?
**(c)** A Fahrenheit degree or a kelvin?

### The Uncertainty of Measurements

**1.28** Define the term *significant figures.*

**1.29** What is meant by rounding a number? Why are numbers rounded? Give a step-by-step process for rounding a number.

**1.30** What is the difference between *accuracy* and *precision*? Explain why one of these two terms is described using *error* and the other is described using *uncertainty*?

**1.31** Suppose a length had been reported to be 31.24 cm. What is the minimum uncertainty implied in this measurement?

**1.32** What is the difference between the treatment of significant figures in addition and multiplication?

### Dimensional Analysis

**1.33** When choosing a conversion factor between two units which are equivalent, there are two possibilities. What is the basic principle when choosing a conversion factor?

**1.34** To convert from mg/mL to mg/L, what conversion factor do you use and why.

**1.35** Suppose someone suggested using the fraction 3 yd/1 ft as a conversion factor to change a length expressed in ft to its equivalent in yds. What is wrong with this conversion factor? Can we construct a valid conversion factor relating cm to m from the equation 1 cm = 1000 m? Explain your answer.

**1.36** In 1 h, there are 3600 s. By what conversion factor would you multiply 250 s to convert it to h? By what conversion factor would you multiply 3.84 h to convert it to seconds?

**1.37** If you were to convert the measured length 4.165 ft to yards by multiplying by the conversion factor (1 yd/3 ft), how many significant figures should the answer contain? Why?

### Density and Specific Gravity

**1.38** Write the equation that defines density. Identify the symbols in the equation.

**1.39** Compare density and specific gravity. What is the difference between the two? When would specific gravity be used?

**1.40** Give four sets of units for density. What mathematical operation must be carried out to convert the density into specific gravity for these four sets of units?

**1.41** Silver has a density of 10.5 g $cm^{-3}$. Express this as an equivalence between mass and volume for silver. Write two conversion factors that can be formed from this equivalence for use in calculations.

# Review Problems

### Physical and Chemical Properties

**1.42** Determine whether each of the following is a physical or chemical change, and explain your reasoning.

**(a)** Copper conducts electricity.
**(b)** Gallium metal melts in your hand.
**(c)** Bread turns brown in a toaster.
**(d)** Wine turns to vinegar.
**(e)** Cement hardens.
**(f)** Baking soda dissolves in water.

**1.43** Determine whether each of the following is a physical or chemical change, and explain your reasoning.

**(a)** Iron rusts.
**(b)** Kernels of corn are heated to make popcorn.
**(c)** Molten copper is mixed with molten gold to make an alloy.
**(d)** Heavy cream is churned to make butter.
**(e)** Water vapor cools into water.
**(f)** Baking soda mixed with vinegar makes bubbles and gases.

**1.44** At room temperature, what is the state of each of the following? If necessary, look up the information in a reference source.

**(a)** hydrogen
**(b)** aluminum
**(c)** nitrogen
**(d)** mercury

**1.45** At room temperature, determine the appropriate phase for each of the following substances. (Look up the substance in data tables if needed.)

**(a)** potassium chloride
**(b)** carbon dioxide
**(c)** methane
**(d)** sucrose

### Measurement of Physical and Chemical Properties

**1.46** What number should replace the question mark in each of the following?

**(a)** 1 zm = ? m
**(b)** 1 km = ? m
**(c)** 1 m = ? pm
**(d)** 1 dm = ? m
**(e)** 1 g = ? kg
**(f)** 1 cg = ? g

**1.47** What numbers should replace the question marks below?

**(a)** 1 nm = ? m
**(b)** 1 μg = ? g
**(c)** 1 kg = ? g
**(d)** 1 Mg = ? g
**(e)** 1 Yg = ? g
**(f)** 1 dg = ? g

**1.48** Perform the following conversions.

**(a)** 57 °C to °F
**(b)** −25.5 °F to °C
**(c)** 378 K to °C
**(d)** −31 °C to K

**1.49** Perform the following conversions.

**(a)** 98 °F to °C
**(b)** −55 °C to °F
**(c)** 299 K to °C
**(d)** 40.0 °C to K

**1.50** The temperature of the core of the sun is estimated to be about 15.7 million K. What is this temperature in °C and °F?

**1.51** Natural gas is mostly methane, a substance that boils at a temperature of 109 K. What is its boiling point in °C and °F?

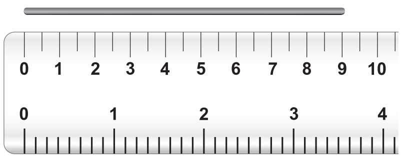
**1.52** A healthy dog has a temperature ranging from 37.8 °C to 39.2 °C. Is a dog with a temperature of 103.5 °F within the normal range?

**1.53** In South America, the warmest recorded temperature of 120.0 °F was in Rivadavia, Argentina, on December 11, 1905. Was Death Valley's record temperature of 56.7 °C on July 10, 1913, warmer or colder than the temperature in South America?

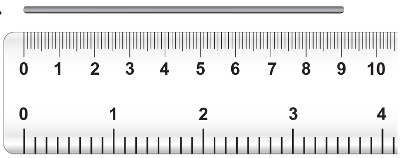
## Uncertainty of Measurements

**1.54** The length of a wire was measured using two different rulers. What are the lengths of the wire and how many significant figures are in each measurement?

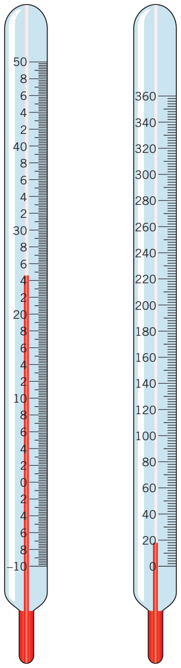
**a.**




**b.**




**1.55** What are the temperatures being measured in the two thermometers, and how many significant figures are in each measurement?



**1.56** How many significant figures do the following measured quantities have?

**(a)** 37.53 cm
**(b)** 37.240 cm
**(c)** 202.0 g
**(d)** 0.00024 kg
**(e)** 2400 mL

**1.57** How many significant figures do the following measured quantities have?

**(a)** 0.0230 g
**(b)** 105.303 m
**(c)** 0.007 kg
**(d)** 614.00 mg
**(e)** 10 L

**1.58** Perform the following mathematical operations and round off the answers to the correct number of significant figures. Include the correct units with the answers.

**(a)** 0.0023 m × 315 m
**(b)** 84.25 kg − 0.01075 kg
**(c)** (184.45 g − 94.45 g)/(31.4 mL − 9.9 mL)
**(d)** (23.4 g + 102.4 g + 0.003 g)/(6.478 mL)
**(e)** (313.44 cm − 209.1 cm) × 8.2234 cm

**1.59** Perform the following mathematical operations and round off the answers to the correct number of significant figures. Include the correct units with the answers.

**(a)** 3.58 g/1.739 mL
**(b)** 4.02 mL + 0.001 mL
**(c)** (22.4 g − 8.3 g)/(1.142 mL − 0.002 mL)
**(d)** (1.345 g − 0.022 g)/(13.36 mL − 8.4115 mL)
**(e)** (74.335 m − 74.332 m)/(4.75 s − 1.114 s)

**1.60** Which are exact numbers and which ones have a finite number of significant figures?

**(a)** the volume of piece of aluminum
**(b)** the number of centimeters in an inch
**(c)** today's temperature
**(d)** the number of seconds in a year

**1.61** Which are exact numbers and which ones have a finite number of significant figures?

**(a)** the mass of 454 g of sugar
**(b)** the distance from the earth to the moon
**(c)** the width of your chemistry book
**(d)** the number of milliliters in a gallon

## Dimensional Analysis

**1.62** Perform the following conversions.

**(a)** 32.0 dm/s to km/hr
**(b)** 8.2 mg/mL to μg/L
**(c)** 75.3 mg to kg
**(d)** 137.5 mL to L
**(e)** 0.025 L to mL
**(f)** 342 $pm^2$ to $dm^2$

**1.63** Perform the following conversions.

**(a)** 92 dL to $mm^3$
**(b)** 22 ng to mg
**(c)** 83 pL to nL
**(d)** 230 $km^3$ to $m^3$
**(e)** 87.3 cm $s^{-2}$ to km $hr^{-2}$
**(f)** 238 $mm^2$ to $nm^2$

**1.64** Perform the following conversions. If necessary, refer to Tables 1.3 and 1.4.

**(a)** 36 in. to cm
**(b)** 5.0 lb to kg
**(c)** 3.0 qt to mL
**(d)** 1 cup (8 oz) to mL
**(e)** 55 mi/hr to km/hr
**(f)** 50.0 mi to km

**1.65** Perform the following conversions. If necessary, refer to Tables 1.3 and 1.4.

**(a)** 250 mL to qt
**(b)** 3.0 ft to m
**(c)** 1.62 kg to lb
**(d)** 1.75 L to fluid oz
**(e)** 35 km/hr to mi/hr
**(f)** 80.0 km to mi

**1.66** Perform the following conversions.

**(a)** 8.4 $ft^2$ to $cm^2$
**(b)** 223 $mi^2$ to $km^2$
**(c)** 231 $ft^3$ to $cm^3$

**1.67** Perform the following conversions.

**(a)** 2.4 $yd^2$ to $m^2$
**(b)** 8.3 $in.^2$ to $mm^2$
**(c)** 9.1 $ft^3$ to L

**1.68** The human stomach can expand to hold up to 4.2 quarts of food. A pistachio nut has a volume of about 0.9 mL. Use this information to estimate the maximum number of pistachios that can be eaten in one sitting.

**1.69** In the movie *Cool Hand Luke* (1967), Luke wagers that he can eat 50 eggs in one hour. The prisoners and guards bet against him, saying, "Fifty eggs gotta weigh a good six pounds. A man's gut can't hold that." A peeled, chewed chicken egg has a volume of approximately 53 mL. If Luke's stomach has a volume of 4.2 qts, does he have any chance of winning the bet?

**1.70** The winds of hurricane Alan in 1980 reached 190. mi/h. What is this speed in m/s? (*Note:* The decimal point after 190 indicates three significant figures.)

**1.71** The U.S. Army's 0.45 caliber Colt bullet used in its M1909 revolver had a muzzle velocity of 738 ft/s. What is this speed expressed in km/h?

**1.72** Battleships from the early 1900's could fire 12-inch diameter artilery shells that had a muzzle velocity of 2700 feet per second. What is this speed in miles per hour?

**1.73** On average, water flows over Niagara Falls at a rate of $2.05 \times 10^5$ $ft^3/s$. One $ft^3$ of water weighs 62.4 lb. Calculate the rate of water flow in tons of water per day. (1 ton = 2000 lb)

**1.74** The brightest star in the night sky in the northern hemisphere is Sirius. Its distance from earth is estimated to be 8.7 light years. A light year is the distance light travels in one year. Light travels at a speed of $3.00 \times 10^8$ m/s. Calculate the distance from earth to Sirius in miles. (1 mi = 5280 ft)

***1.75** One degree of latitude on the earth's surface equals 60.0 nautical miles. One nautical mile equals 1.151 statute miles. (A *statute mile* is the distance over land that we normally associate with the unit mile.) Calculate the circumference of the earth in statute miles.

### Density and Specific Gravity

**1.76** Calculate the density of kerosene, in g/mL, if its mass is 36.4 g and its volume is 45.6 mL.

**1.77** Calculate the density of magnesium, in g/$cm^3$, if its mass is 14.3 g and its volume is 8.46 $cm^3$.

**1.78** Acetone, the solvent in some nail polish removers, has a density of 0.791 g/mL. What is the volume, in mL, of 25.0 g of acetone?

**1.79** A glass apparatus, called a pycnometer, contains 26.223 g of water when filled at 25 °C. At this temperature, water has a density of 0.99704 g/mL. What is the volume, in mL, of the apparatus?

**1.80** Chloroform, a chemical once used as an anesthetic, has a density of 1.492 g/mL. What is the mass in grams of 185 mL of chloroform?

**1.81** Gasoline's density is about 0.65 g/mL. How much does 34 L (approximately 18 gals) weigh in kg? In lbs?

**1.82** A graduated cylinder was filled with water to the 15.0 mL mark and weighed on a balance. Its mass was 27.35 g. An object made of silver was placed in the cylinder and completely submerged in the water. The water level rose to 18.3 mL. When reweighed, the cylinder, water, and silver object had a total mass of 62.00 g. Calculate the density of silver in g $cm^{-3}$.

**1.83** Titanium is a metal used to make golf clubs. A rectangular bar of this metal measuring 1.84 cm × 2.24 cm × 2.44 cm was found to have a mass of 45.7 g. What is the density of titanium in g $cm^{-3}$?

**1.84** The space shuttle uses liquid hydrogen as its fuel. The external fuel tank used during takeoff carries 227,641 lb of hydrogen with a volume of 385,265 gal. Calculate the density of liquid hydrogen in units of lb/gal and g/mL. (Express your answer to three significant figures.) What is the specific gravity of liquid hydrogen?

**1.85** You are planning to make a cement driveway that has to be 10.1 ft wide, 32.3 ft long, and 4.00 in. deep on average. The specific gravity of concrete is 0.686. How many kg of concrete are needed for the job?

## Additional Exercises

**1.86** Some time ago, a U.S. citizen traveling in Canada observed that the price of regular gasoline was 1.428 Canadian dollars/L. The exchange rate at the time was 1.247 Canadian dollars per one U.S. dollar. Calculate the price of the Canadian gasoline in units of U.S. dollars/gal. (Just the week before, the traveler had paid $3.759 per gal in the United States.)

**1.87** Driving to work one day, one of the authors of this book had a gas mileage of 28.5 km/L. Another author got 31 mi/gal. Which author had the more efficient car? Why?

**1.88** In lab, you are given two metal rods *A* and *B*. One is made of zinc (density = 7.14 g/$cm^3$), and the other is made of nickel (density = 8.90 g/$cm^3$), and you are asked to determine which one is which. Using a graduated cylinder and a balance, you find that the volume of the rod *A* is 4.2 $cm^3$ and the volume of rod *B* is 3.7 $cm^3$. Weighing the two rods, the mass of rod *A* is 37.8 g and the mass of rod *B* is 26.2 g. Which rod is which metal?

**1.89** You are the science reporter for a daily newspaper, and your editor has asked you to write a story based on a report in the scientific literature. The report states that analysis of the sediments in Hausberg Tarn (elevation 4350 m) on the side of Mount Kenya (elevation 4600–4700 m) shows that the average temperature of the water rose by 4.0 °C between 350 BC and 450 AD. Your editor wants all the data expressed in the English system of units. Make the appropriate conversions.

**1.90** An astronomy web site states that neutron stars have a density of $1.00 \times 10^8$ tons per cubic centimeter. The site does not specify whether "tons" means metric tons (1 metric ton = 1000 kg) or English tons (1 English ton = 2000 pounds). How many grams would one teaspoon of a neutron star weigh if the density were in metric tons per $cm^3$? How many grams would the teaspoon weigh if the density were in English tons per $cm^3$? (One teaspoon is defined as 5.00 mL.)

**1.91** The star Arcturus is $3.50 \times 10^{14}$ km from the earth. How many days does it take for light to travel from Arcturus to earth? What is the distance to Arcturus in light years? One light year is the distance light travels in one year (365 days); light travels at a speed of $3.00 \times 10^8$ m/s.

**1.92** A pycnometer is a glass apparatus used for accurately determining the density of a liquid. When dry and empty, a certain pycnometer had a mass of 27.314 g. When filled with distilled water at 25.0 °C, it weighed 36.842 g. When filled with chloroform (a liquid once used as an anesthetic before its toxic properties were known), the apparatus weighed 41.428 g. At 25.0 °C, the density of water is 0.99704 g/mL. **(a)** What is the volume of the pycnometer? **(b)** What is the density of chloroform?

**1.93** Radio waves travel at the speed of light, $3.00 \times 10^8$ m/s. If you were to broadcast a question to the Mars Rover, Spirit, on Mars, which averages 225,000,000 miles from earth, what is the minimum time that you would have to wait to receive a reply?

**1.94** Suppose you have a job in which you earn $7.35 for each 30 minutes that you work.

**(a)** Express this information in the form of an equivalence between dollars earned and minutes worked.

**(b)** Use the equivalence defined in **(a)** to calculate the number of dollars earned in 1 hr 45 min.

**(c)** Use the equivalence defined in **(a)** to calculate the number of minutes you would have to work to earn $333.50.

**1.95** When an object floats in water, it displaces a volume of water that has a weight equal to the weight of the object, Archimedes principle. If a ship has a weight of 4255 tons, how many cubic feet of seawater will it displace? Seawater has a density of 1.025 g $cm^{-3}$; 1 ton = 2000 lb, exactly.

**1.96** Aerogel or "solid smoke" is a novel material that is made of silicon dioxide, like glass, but is a thousand times less dense than glass because it is extremely porous. Material scientists at NASA's Jet Propulsion Laboratory created the lightest aerogel ever in 2002, with a density of 0.00011 pounds per cubic inch. The material was used for thermal insulation in the 2003 Mars Exploration Rover. If the maximum space for insulation in the spacecraft's hull was 2510 $cm^3$, what mass (in grams) did the aerogel insulation add to the spacecraft?


NASA/JPL

**1.97** A liquid known to be either ethanol (ethyl alcohol) or methanol (methyl alcohol) was found to have a density of $0.798 \pm 0.001$ g/mL. Consult tabulated data to determine which liquid it is. What other measurements could help to confirm the identity of the liquid?

**1.98** What would be the volume of a piece of osmium metal if it had the same mass as a piece of tin with a volume of 9.25 $cm^3$.

**1.99** An unknown liquid was found to have a density of 69.22 $lb/ft^3$. The density of ethylene glycol (the liquid used in antifreeze) is 1.1088 g/mL. Could the unknown liquid be ethylene glycol?

**1.100** For diamonds, 1.00 carat is 0.200 g, and the density of diamond is 3.51 $g/cm^3$. What is the volume of a 1.75 carat diamond?

**1.101** There exists a single temperature at which the value reported in °F is numerically the same as the value reported in °C. What is this temperature?

**1.102** In the text, the Kelvin scale of temperature is defined as an absolute scale in which one Kelvin degree unit is the same size as one Celsius degree unit. A second absolute temperature scale exists called the Rankine scale. On this scale, one Rankine degree unit (°R) is the same size as one Fahrenheit degree unit. **(a)** What is the only temperature at which the Kelvin and Rankine scales possess the same numerical value? Explain your answer. **(b)** What is the boiling point of water expressed in °R?

***1.103** Density measurements can be used to analyze mixtures. For example, the density of solid sand (without air spaces) is about 2.84 g/mL. The density of gold is 19.3 g/mL. If a 1.00 kg sample of sand containing some gold has a density of 3.10 g/mL (without air spaces), what is the percentage of gold in the sample?

***1.104** An artist's statue has a surface area of 14.6 $ft^2$. The artist plans to apply gold leaf to the statue and wants the coating to be 2.50 μm thick. If the price of gold was $1910.50 per troy ounce on 10/13/2020, how much would it cost to give the statue its gold coating? (1 troy ounce = 31.1035 g; the density of gold is 19.3 $g/cm^3$.)

**1.105** What is the volume in cubic millimeters of a 3.54 carat diamond, given that the density of diamond is 3.51 g/mL (1 carat = 200 mg)? Assuming that the diamond is generally spherical, what is the approximate diameter of the diamond in mm?

# Exercises in Critical Thinking

**1.106** A solution is defined as a uniform mixture consisting of a single phase. With our vastly improved abilities to "see" smaller and smaller particles, down to the atomic level, present an argument for the proposition that all mixtures are heterogeneous. Present the argument that the ability to observe objects as small as an atom has no effect on the definitions of heterogeneous and homogeneous.

**1.107** How do you know that Coca-Cola is not a compound? What experiments could you perform to prove it? What would you do to prove that the rusting of iron is a chemical change rather than a physical change?

**1.108** Find two or more web sites that give the values for each of the seven base SI units. Keeping in mind that not all web sites provide reliable information, which web site do you believe provides the most reliable values? Justify your answer.

**1.109** Reference books such as the *Handbook of Chemistry and Physics* report the specific gravities of substances instead of their densities. Using the definition of specific gravity, discuss the relative merits of specific gravity and density in terms of their usefulness as a physical property.

***1.110** A student used a 250 mL graduated cylinder having volume markings every 2 mL to carefully measure 100 mL of water for an experiment. A fellow student said that by reporting the volume as "100 mL" in her lab notebook, she was only entitled to one significant figure. The first student disagreed. Why did her fellow student say the reported volume had only one significant figure?

***1.111** Because of the serious consequences of lead poisoning, the Federal Centers for Disease Control in Atlanta changed the "a threshold of concern" for lead levels in children's blood in 2012. This threshold was based on a study that suggested that lead levels in blood as low as *5 micrograms of lead per deciliter of blood* can result in subtle effects of lead toxicity in children up to 5 years old. Suppose a child had a lead level in her blood of $2.5 \times 10^{-4}$ grams of lead per liter of blood. Is this person in danger of exhibiting the effects of lead poisoning?

***1.112** Gold has a density of 19.31 g $cm^{-3}$. How many grams of gold are required to provide a gold coating 0.500 mm thick on a ball bearing having a diameter of 2.000 mm?

***1.113** A Boeing 747 jet airliner carrying 568 people burns about 5.0 gallons of jet fuel per mile. What is the rate of fuel consumption in units of gallons per person per mile? Is this better or worse than the rate of fuel consumption in an automobile carrying two people that gets 21.5 miles per gallon? If the airliner were making the 3470 mile trip from New York to London, how many pounds of jet fuel would be consumed? (Jet fuel has a density of 0.817 g/mL.) Considering the circumstances, how many significant figures should be used in reporting the volume of jet fuel consumed? Justify your answer.

***1.114** Download a table of data for the density of water between its freezing and boiling points. Use a spreadsheet program to plot **(a)** the density of water versus temperature and **(b)** the volume of a kilogram of water versus temperature. Interpret the significance of these plots.

**1.115** Imagine a new temperature scale based on butanol where the boiling point of butanol is 117.7 °C and the freezing point of butanol is −89.8 °C. In this new scale, with the symbol °B, the 0 °B is the freezing point and the 100 °B is the boiling point of butanol. What would the boiling point of water be? What would room temperature be, that is 25 °C?

***1.116** A barge pulled by a tugboat can be described as a rectangular box that is open on the top. A certain barge is 30.2 feet wide and 12.50 feet high, with the base of each side being 116 feet long. If the barge itself weighs $6.58 \times 10^4$ pounds, what will be the draft (depth in the water) if the cargo weighs $1.12 \times 10^6$ pounds in **(a)** seawater with a density of 1.025 g/mL and **(b)** freshwater with a density of 0.997 g/mL?

**1.117** Pepsi Cola™ contains several substances from all three states of matter. Name at least one substance from each of the three states. (consult the label on the bottle or can). When your Pepsi goes flat, is that a chemical or physical change?

**1.118** Intensive properties stay constant when the size of the sample changes, while extensive properties change with the size of the sample. Why are both intensive and extensive properties important for scientists?