Matter and Measurements

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Museum of the History of Science in Oxford, England

The painting shows measuring instruments used in the Middle Ages. We still use many of them today.

There is measure in everything.

—HORACE

Almost certainly, this is your first college course in chemistry; perhaps it is your first exposure to chemistry at any level. Unless you are a chemistry major, you may wonder why you are taking this course and what you can expect to gain from it. To address that question, it is helpful to look at some of the ways in which chemistry contributes to other disciplines.

If you're planning to be an engineer, you can be sure that many of the materials you will work with have been synthesized by chemists. Some of these materials are organic (carbon-containing). They could be familiar plastics like polyethylene (Chapter 23) or the more esoteric plastics used in unbreakable windows and nonflammable clothing. Other materials, including metals (Chapter 20) and semiconductors, are inorganic in nature.

Perhaps you are a health science major, looking forward to a career in medicine or pharmacy. If so, you will want to become familiar with the properties of aqueous solutions (Chapters 4, 10, 14, and 16), which include blood and other body fluids. Chemists today are involved in the synthesis of a variety of life-saving products. These range from drugs used in chemotherapy (Chapter 19) to new antibiotics used against resistant microorganisms.

Beyond career preparation, an objective of a college education is to make you a better-informed citizen. In this text, we'll look at some of the chemistry-related topics that make the news:

- depletion of the ozone layer (Chapter 11).
- alternative sources of fuel (Chapter 17).
- the pros and cons of nuclear power (Chapter 18).

Another goal of this text is to pique your intellectual curiosity by trying to explain the chemical principles behind such recent advances as

- "self-cleaning" windows (Chapter 1).
- "the ice that burns" (Chapter 3).
- "maintenance-free" storage batteries (Chapter 17).
- "chiral" drugs (Chapter 22).

We hope that when you complete this course you too will be convinced of the importance of chemistry in today's world. We should, however, caution you on one point.

Chapter Outline

Chemistry deals with the properties and reactions of substances.

Although we will talk about many of the applications of chemistry, *our main concern will be with the principles that govern chemical reactions.* Only by mastering those principles will you understand the basis of the applications mentioned above.

This chapter begins the study of chemistry by

- considering the different types of matter: pure substances versus mixtures, elements versus compounds (Section 1-1).
- looking at the kinds of measurements fundamental to chemistry, the uncertainties associated with those measurements, and a method to convert measured quantities from one unit to another (Section 1-2).
- focusing on certain physical properties, including density and water solubility, which can be used to identify substances (Section 1-3).

1-1 Matter and Its Classifications

Matter is anything that has mass and occupies space. It can be classified either with respect to its physical phases or with respect to its composition (Figure 1.1).

The three phases of matter are solid, liquid, and gas. A **solid** has a fixed shape and volume. A **liquid** has a fixed volume but is not rigid in shape; it takes the shape of its container. A **gas** has neither a fixed volume nor a shape. It takes on both the shape and the volume of its container.

Matter can also be classified with respect to its composition:

- pure substances, each of which has a fixed composition and a unique set of properties.
- mixtures, composed of two or more substances.

Pure substances are either elements or compounds (Figure 1.1), whereas mixtures can be either homogeneous or heterogeneous.

Elements

An **element** is a type of matter that cannot be broken down into two or more pure substances. There are 118 known elements, of which 91 occur naturally.

Many elements are familiar to all of us. The charcoal used in outdoor grills is nearly pure carbon. Electrical wiring, jewelry, and water pipes are often made from copper, a metallic element. Another such element, aluminum, is used in many household utensils.

Some elements come in and out of fashion, so to speak. Sixty years ago, elemental silicon was a chemical curiosity. Today, ultrapure silicon has become the basis for the multibillion-dollar semiconductor industry. Lead, on the other hand, is an element moving in the other direction. A generation ago it was widely used to make paint pigments, plumbing connections, and gasoline additives. Today, because of the toxicity of lead compounds, all of these applications have been banned in the United States.

In chemistry, an element is identified by its **symbol**. This consists of one or two letters, usually derived from the name of the element. Thus the symbol for carbon is C; that for aluminum is Al. Sometimes the symbol comes from the Latin name of the element or one of its compounds. The two elements copper and mercury, which were known in ancient times, have the symbols Cu *(cuprum)* and Hg *(hydrargyrum)*.

Table 1.1 (p. 4) lists the names and symbols of several elements that are probably familiar to you. In either free or combined form, they are commonly found in the laboratory or in commercial products. The abundances listed measure the relative amount of each element in the earth's crust, the atmosphere, and the oceans.

Curiously, several of the most familiar elements are really quite rare. An example is mercury, which has been known since at least 500 B.C., even though its abundance is only 0.00005%. It can easily be prepared by heating the red mineral cinnabar (Figure 1.2, p. 4).

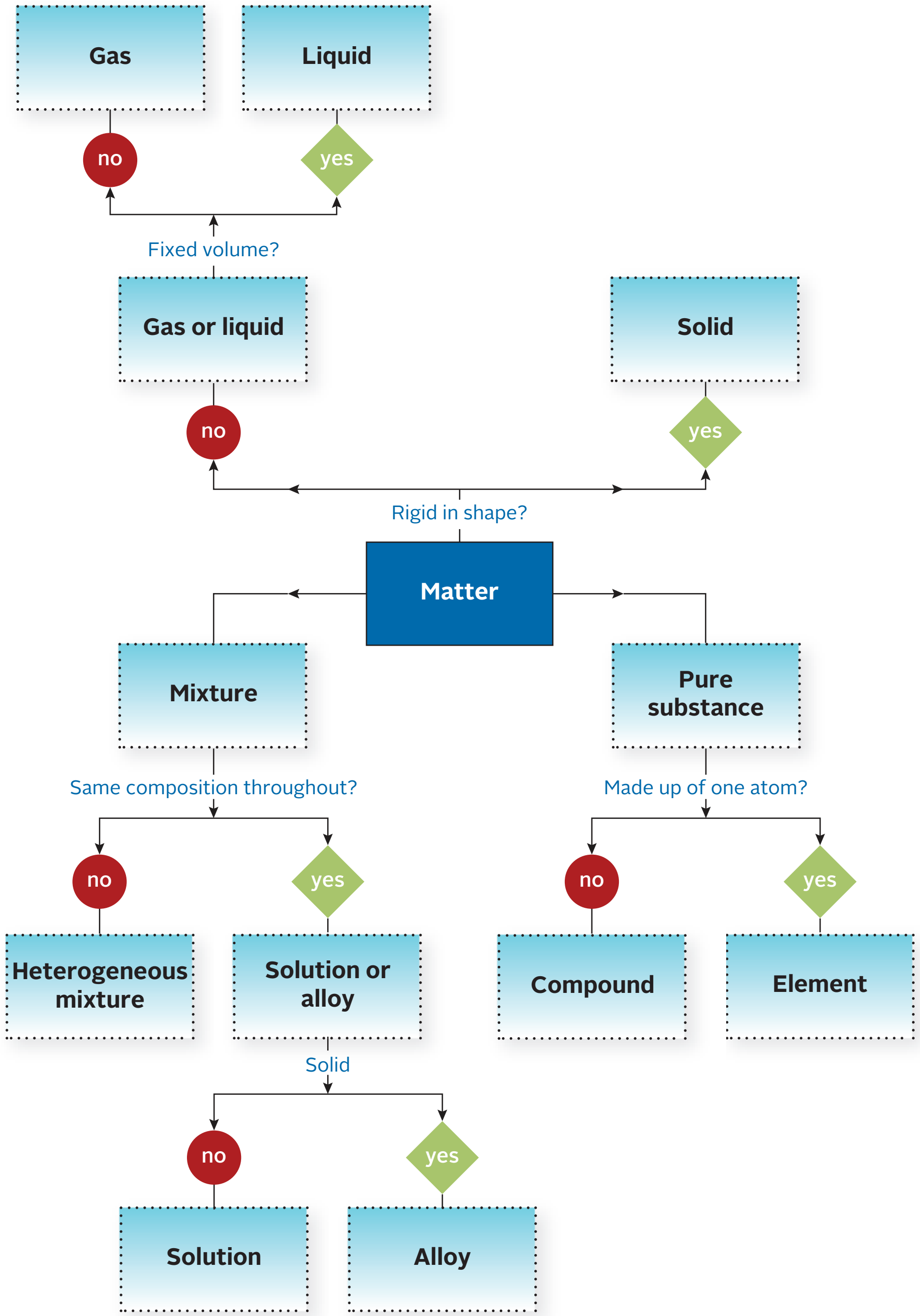


Figure 1.1 **Classification of matter into solid, liquid, and gas.**

Mercury is the only metal that is a liquid at room temperature. It is also one of the densest elements. Because of its high density, mercury was the liquid extensively used in thermometers and barometers. In the 1990s all instruments using mercury were banned because of environmental concerns. Another useful quality of mercury is its ability to dissolve many metals, forming solutions (amalgams). A silver-mercury-tin amalgam is still used to fill tooth cavities, but many dentists now use tooth-colored composites because they adhere better and are aesthetically more pleasing.

Mercury thermometers, both for laboratory and clinical use, have been replaced by digital ones.

In contrast, aluminum (abundance = 7.5%), despite its usefulness, was little more than a chemical curiosity until about a century ago. It occurs in combined form in clays and rocks, from which it cannot be extracted. In 1886 two young chemists, Charles Hall in the United States and Paul Héroult in France, independently worked out a process for extracting aluminum from a relatively rare ore,

Table 1.1 Some Familiar Elements with Their Percentage Abundances

Element	Symbol	Percentage Abundance	Element	Symbol	Percentage Abundance
Aluminum	Al	7.5	Manganese	Mn	0.09
Bromine	Br	0.00025	Mercury	Hg	0.00005
Calcium	Ca	3.4	Nickel	Ni	0.010
Carbon	C	0.08	Nitrogen	N	0.03
Chlorine	Cl	0.2	Oxygen	O	49.4
Chromium	Cr	0.018	Phosphorus	P	0.12
Copper	Cu	0.007	Potassium	K	2.4
Gold	Au	0.0000005	Silicon	Si	25.8
Hydrogen	H	0.9	Silver	Ag	0.00001
Iodine	I	0.00003	Sodium	Na	2.6
Iron	Fe	4.7	Sulfur	S	0.06
Lead	Pb	0.0016	Titanium	Ti	0.56
Magnesium	Mg	1.9	Zinc	Zn	0.008

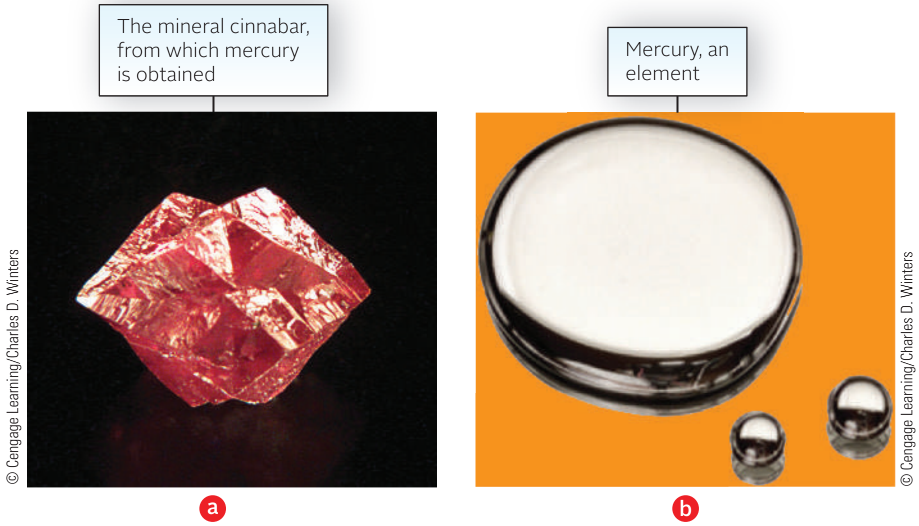


Figure 1.2 **Cinnabar and mercury.**

bauxite. That process is still used today to produce the element. By an odd coincidence, Hall and Héroult were born in the same year (1863) and died in the same year (1914).

Compounds

A **compound** is a pure substance that contains more than one element. Water is a compound of hydrogen and oxygen. The compounds methane, acetylene, and naphthalene all contain the elements carbon and hydrogen, in different proportions.

Compounds have fixed compositions. That is, a given compound always contains the same elements in the same percentages by mass. A sample of pure water contains precisely 11.19% hydrogen and 88.81% oxygen. In contrast, mixtures can vary in composition. For example, a mixture of hydrogen and oxygen might contain 5, 10, 25, or 60% hydrogen, along with 95, 90, 75, or 40% oxygen.

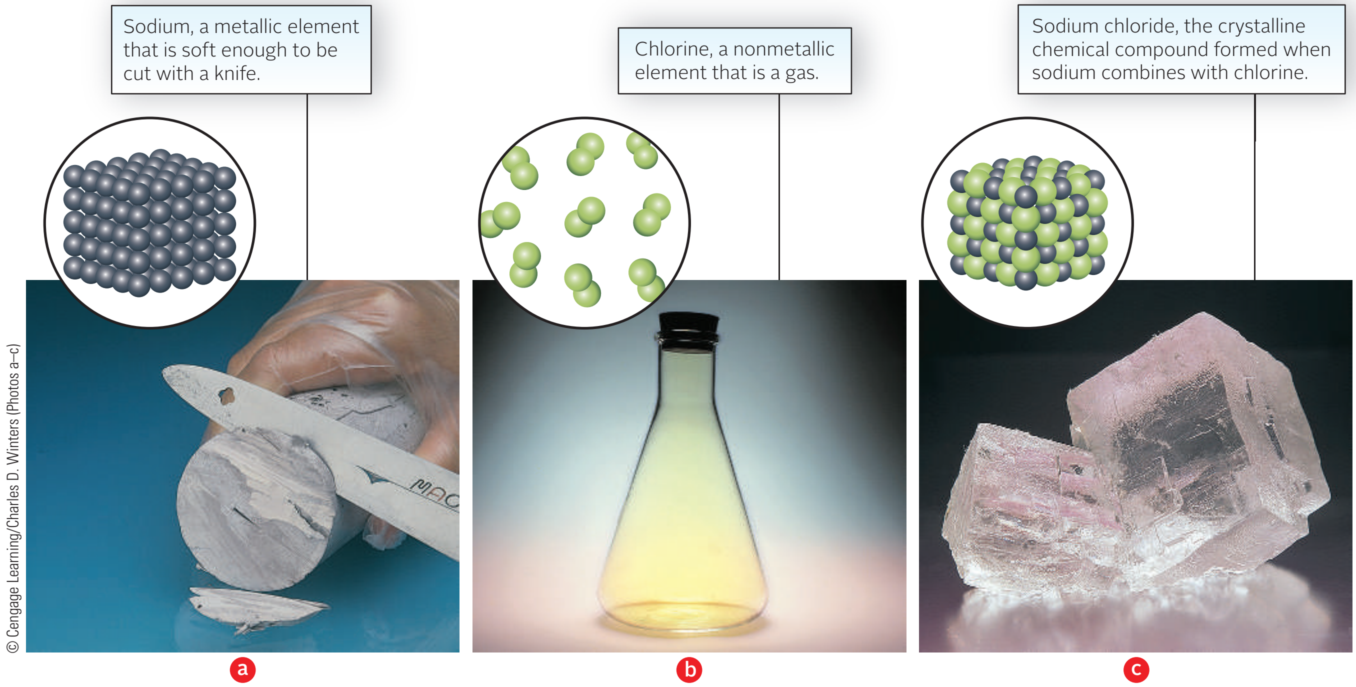


Figure 1.3 **Sodium, chlorine, and sodium chloride.**

The properties of compounds are usually very different from those of the elements they contain. Ordinary table salt, sodium chloride, is a white, unreactive solid. As you can guess from its name, it contains the two elements sodium and chlorine. Sodium (Na) is a shiny, extremely reactive metal. Chlorine (Cl) is a poisonous, greenish-yellow gas. Clearly, when these two elements combine to form sodium chloride, a profound change takes place (Figure 1.3).

Many different methods can be used to resolve compounds into their elements. Sometimes, but not often, heat alone is sufficient. Mercury(II) oxide, a compound of mercury and oxygen, decomposes to its elements when heated to 600°C. Joseph Priestley, an English chemist, discovered oxygen more than 200 years ago when he carried out this reaction by exposing a sample of mercury(II) oxide to an intense beam of sunlight focused through a powerful lens. The mercury vapor formed is a deadly poison. Sir Isaac Newton, who distilled large quantities of mercury in his laboratory, suffered the effects in his later years.

Another method of resolving compounds into elements is *electrolysis*, which involves passing an electric current through a compound, usually in the liquid state. By electrolysis it is possible to separate water into the gaseous elements hydrogen and oxygen. Several decades ago it was proposed to use the hydrogen produced by electrolysis to raise the *Titanic* from its watery grave off the coast of Newfoundland. It didn't work.

Figure 1.4 **A heterogeneous mixture of copper sulfate crystals (blue) and sand.**

Mixtures

A **mixture** contains two or more substances combined in such a way that each substance retains its chemical identity. When you shake copper sulfate with sand (Figure 1.4), the two substances do not react with one another. In contrast, when sodium is exposed to chlorine gas, a new compound, sodium chloride, is formed.

Most materials you encounter are mixtures.

There are two types of mixtures:

1. Homogeneous or uniform mixtures are ones in which the composition is the same throughout. Another name for a homogeneous mixture is a **solution,** which is made up of a solvent, usually taken to be the substance present in largest amount, and one or more solutes. Most commonly, the **solvent** is a liquid, whereas

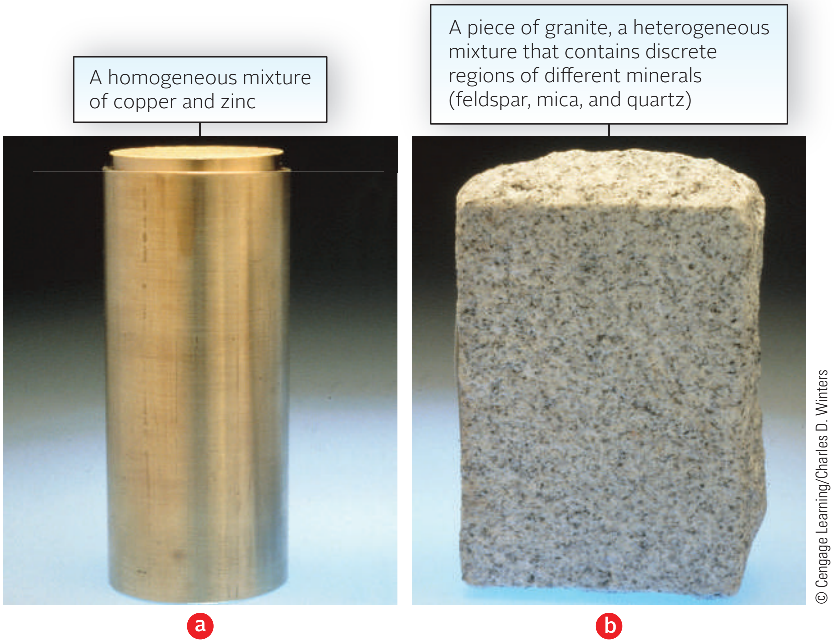


Figure 1.5 **Two mixtures.**

a **solute** may be a solid, liquid, or gas. Soda water is a solution of carbon dioxide (solute) in water (solvent). Seawater is a more complex solution in which there are several solid solutes, including sodium chloride; the solvent is water. It is also possible to have solutions in the solid state. Brass (Figure 1.5a) is a solid solution containing the two metals copper (67%–90%) and zinc (10%–33%).

All gaseous mixtures, including air, are solutions.

2. **Heterogeneous** or nonuniform mixtures are those in which the composition varies throughout. Most rocks fall into this category. In a piece of granite (Figure 1.5b), several components can be distinguished, differing from one another in color.

Many different methods can be used to separate the components of a mixture from one another. A couple of methods that you may have carried out in the laboratory are

- **filtration**, used to separate a heterogeneous solid-liquid mixture. The mixture is passed through a barrier with fine pores, such as filter paper. Copper sulfate, which is water-soluble, can be separated from sand by shaking with water. On filtration the sand remains on the paper and the copper sulfate solution passes through it.
- *distillation,* used to resolve a homogeneous solid-liquid mixture. The liquid vaporizes, leaving a residue of the solid in the distilling flask. The liquid is obtained by condensing the vapor. Distillation can be used to separate the components of a water solution of copper sulfate (Figure 1.6).

A more complex but more versatile separation method is *chromatography,* a technique widely used in teaching, research, and industrial laboratories to separate all kinds of mixtures. This method takes advantage of differences in solubility and/or extent of adsorption on a solid surface. In *gas-liquid chromatography,* a mixture of volatile liquids and gases is introduced into one end of a heated glass tube. As little as one microliter (10^{-6} L) of sample may be used. The tube is packed with an inert solid whose surface is coated with a viscous liquid. An unreactive "carrier gas," often helium, is passed through the tube. The components of the sample gradually separate as they vaporize into the helium or condense into the viscous liquid. Usually the more volatile fractions move faster and emerge first; successive fractions activate a detector and recorder.

Gas-liquid chromatography (GLC) (Figure 1.7) finds many applications outside the chemistry laboratory. If you've ever had an emissions test on the exhaust system of your car, GLC was almost certainly the analytical method used. Pollutants such as carbon monoxide and unburned hydrocarbons appear as peaks on a graph. A computer determines the areas under these peaks, which are proportional to the concentrations of pollutants, and prints out a series of numbers that tells the inspector whether your car passed or failed the test. Many of the techniques used to test people for drugs (marijuana, cocaine, and others) or alcohol also make use of gas-liquid chromatography.

GLC is a favorite technique in the forensics labs of many TV shows.

Ultra-high-speed gas chromatography (GC) fitted with an odor sensor is a powerful tool for analyzing the chemical vapors produced by explosives or other chemical or biological weapons.

In this section we will look at four familiar properties that you will almost certainly measure in the laboratory: length, volume, mass, and temperature. Other physical and chemical properties will be introduced in later chapters as they are needed.

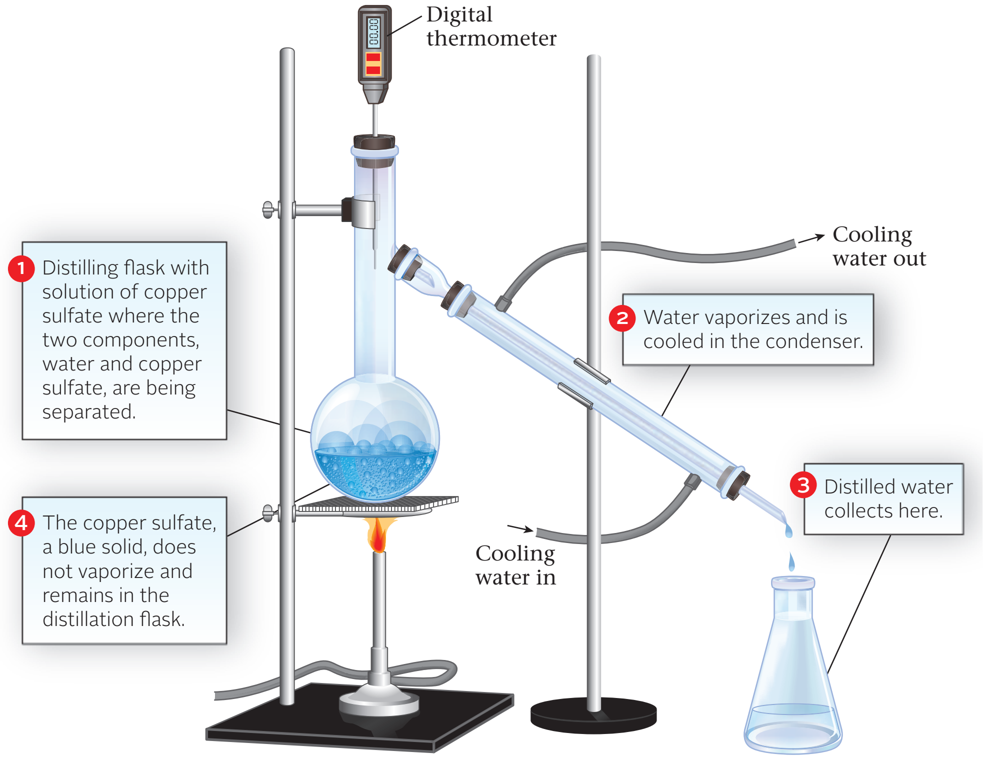


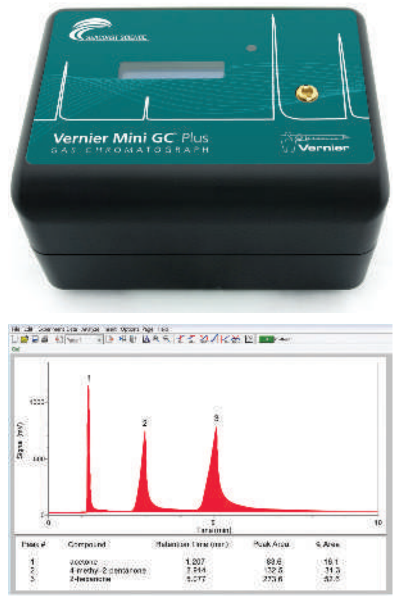
Figure 1.6 Apparatus for a simple distillation.

1-2 Measurements

Chemistry is a quantitative science. The experiments that you carry out in the laboratory and the calculations that you perform almost always involve measured quantities with specified numerical values. Consider, for example, the following set of directions for the preparation of aspirin (measured quantities are shown in italics).

> Add *2.0 g* of salicylic acid, *5.0 mL* of acetic anhydride, and *5 drops* of *85%* H_3PO_4 to a 50-mL Erlenmeyer flask. Heat in a water bath at *75°C* for *15 minutes*. Add cautiously *20 mL* of water and transfer to an ice bath at *0°C*. Scratch the inside of the flask with a stirring rod to initiate crystallization. Separate aspirin from the solid-liquid mixture by filtering through a Buchner funnel *10 cm* in diameter.

Scientific measurements are expressed in the **metric system**. As you know, this is a decimal-based system in which all of the units of a particular quantity are related to one another by factors of 10. The more common prefixes used to express these factors are listed in Table 1.2.



Vernier

Vernier

Figure 1.7 Mass spectrometer and gas chromatograph. Airport security uses these instruments to separate mixtures (chromatograph) and to detect the presence of nitrogen containing explosives (mass spectrograph). Larger versions of these instruments are used to scan checked baggage.

Table 1.2 Metric Prefixes

Factor	Prefix	Abbreviation	Factor	Prefix	Abbreviation
10^6	mega	M	10^{-3}	**milli**	m
10^3	**kilo**	k	10^{-6}	micro	μ
10^{-1}	deci	d	10^{-9}	**nano**	n
10^{-2}	**centi**	c	10^{-12}	pico	p

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Figure 1.8 Measuring volume. A buret (*left*) delivers an accurately measured variable volume of liquid. A pipet (*right*) delivers a fixed volume (e.g., 25.00 mL) of liquid.

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Figure 1.9 Weighing a solid. The solid sample plus the paper on which it rests weighs 144.998 g. The pictured balance is a single-pan analytical balance.

Writing "m" in upper case or lower case makes a big difference.

Many countries still use degrees centigrade.

Instruments and Units

The standard unit of *length* in the metric system is the meter, which is a little larger than a yard. The meter was originally intended to be 1/40,000,000 of the earth's meridian that passes through Paris. It is now defined as the distance light travels in a vacuum in 1/299,792,458 of a second.

Other units of length are expressed in terms of the meter, using the prefixes listed in Table 1.2. You are familiar with the centimeter, the millimeter, and the kilometer:

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

The dimensions of very tiny particles are often expressed in nanometers:

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

Volume is most commonly expressed in one of three units

- cubic centimeters $1\text{ cm}^3 = (10^{-2}\text{ m})^3 = 10^{-6}\text{ m}^3$
- liters (L) $1\text{ L} = 10^{-3}\text{ m}^3 = 10^{3}\text{ cm}^3$
- milliliters (mL) $1\text{ mL} = 10^{-3}\text{ L} = 10^{-6}\text{ m}^3$

Notice that a milliliter is equal to one cubic centimeter:

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

The device most commonly used to measure volume in general chemistry is the graduated cylinder. A pipet or buret (Figure 1.8) is used when greater accuracy is required. A pipet is calibrated to deliver a fixed volume of liquid—for example, 25.00 mL—when filled to the mark and allowed to drain. Different volumes can be delivered accurately by a buret, perhaps to ±0.01 mL.

In the metric system, *mass* is most commonly expressed in grams, kilograms, or milligrams:

$$1\text{ g} = 10^{-3}\text{ kg} \qquad 1\text{ mg} = 10^{-3}\text{ g}$$

This book weighs about 1.5 kg. The megagram, more frequently called the *metric ton,* is

$$1\text{ Mg} = 10^{6}\text{ g} = 10^{3}\text{ kg}$$

Properly speaking, there is a distinction between mass and weight. *Mass* is a measure of the amount of matter in an object; *weight* is a measure of the gravitational force acting on the object. Chemists often use these terms interchangeably; we determine the mass of an object by "weighing" it on a balance (Figure 1.9).

Temperature is the factor that determines the direction of heat flow. When two objects at different temperatures are placed in contact with one another, heat flows from the one at the higher temperature to the one at the lower temperature.

Thermometers used in chemistry are marked in degrees **Celsius** (referred to as degrees centigrade until 1948). On this scale, named after the Swedish astronomer Anders Celsius (1701–1744), the freezing point of water is taken to be 0°C. The normal boiling point of water is 100°C. Household thermometers in the United States are commonly marked in *Fahrenheit* degrees. Daniel Fahrenheit (1686–1736) was a German instrument maker who was the first to use the mercury-in-glass thermometer. On this scale, the normal freezing and boiling points of water are taken to be 32° and 212°, respectively (Figure 1.10). It follows that (212°F − 32°F) = 180°F covers the same temperature interval as (100°C − 0°C) = 100°C. This leads to the general relation between the two scales:

$$t_{°F} = 1.8\, t_{°C} + 32° \tag{1.1}$$

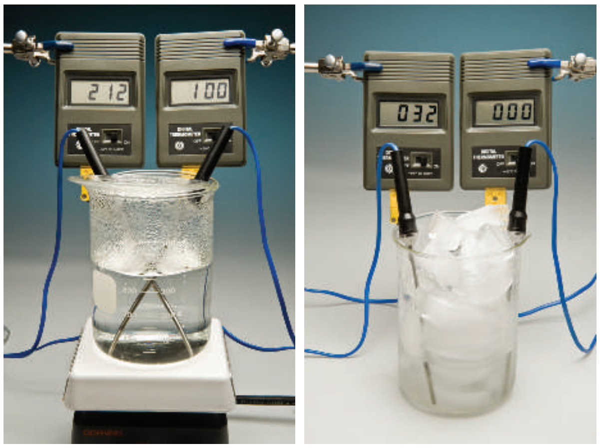

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Figure 1.10 Relationship between Fahrenheit and Celsius scales. This figure shows the relationship between the Fahrenheit and Celsius temperature scales. Note that there are 180 degrees F for 100 degrees C (1.8 F/C) and 0°C = 32°F.

The two scales coincide at $-40°$; as you can readily see from equation 1.1:

$$\text{At } -40°\text{C:} \qquad t_{°F} = 1.8(-40°) + 32° = -72° + 32° = -40°$$

For many purposes in chemistry, the most convenient unit of temperature is the **kelvin** (K); note the absence of the degree sign. The kelvin is defined to be 1/273.16 of the difference between the lowest attainable temperature (0 K) and the triple point of water* (0.01°C). The relationship between temperature in K and in °C is

$$T_K = t_{°C} + 273.15 \qquad \textbf{(1.2)}$$

This scale is named after Lord Kelvin (1824–1907), a British scientist who showed in 1848, at the age of 24, that it is impossible to reach a temperature lower than 0 K.

EXAMPLE 1.1

Mercury thermometers have been phased out because of the toxicity of mercury vapor. A common replacement for mercury in glass thermometers is the organic liquid isoamyl benzoate, which boils at 262°C. What is its boiling point in (a) °F? (b) K?

ANALYSIS	
Information given:	Boiling point (262°C)
Asked for:	boiling point in °F and K

STRATEGY

1. Substitute into Equation 1.1 for (a).

2. Substitute into Equation 1.2 for (b).

SOLUTION	
(a) °F	°F = 1.8(°C) + 32 = 1.8(262°C) + 32 = **504°F**
(b) K	K = 273.15 + 262°C = **535 K**

As you can see from this discussion, a wide number of different units can be used to express measured quantities in the metric system. This proliferation of units has long been of concern to scientists. In 1960 a self-consistent set of metric units was proposed. This so-called International System of Units (SI) is discussed in Appendix 1. The SI units for the four properties we have discussed so far are

Length:	meter (m)	*Mass:*	kilogram (kg)
Volume:	cubic meter (m^3)	*Temperature:*	kelvin (K)

Uncertainties in Measurements: Significant Figures

Every measurement carries with it a degree of uncertainty. Its magnitude depends on the nature of the measuring device and the skill of its operator. Suppose, for example, you measure out 8 mL of liquid using the 100-mL graduated cylinder shown in Figure 1.11. Here the volume is uncertain to perhaps ±1 mL. With such a crude measuring device, you would be lucky to obtain a volume between 7 and 9 mL. To obtain greater precision, you could use a narrow 10-mL cylinder, which has divisions in small increments. You might now measure a volume within 0.1 mL of the desired value, in the range of 7.9 to 8.1 mL. By using a buret, you could reduce the uncertainty to ±0.01 mL.

Anyone making a measurement has a responsibility to indicate the uncertainty associated with it. Such information is vital to someone who wants to repeat the

*The triple point of water (Chapter 9) is the one unique temperature and pressure pair at which ice, liquid water, and water vapor can coexist in contact with one another.

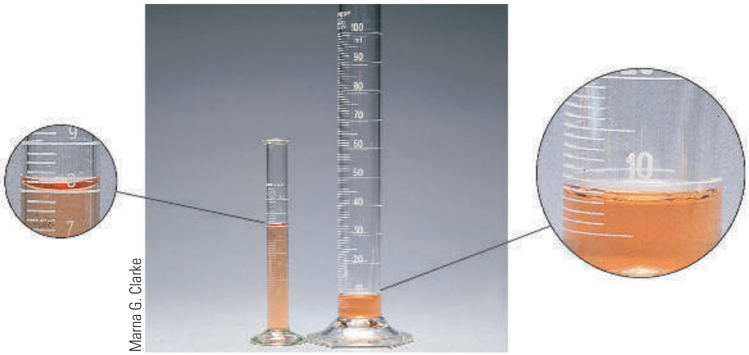

Figure 1.11 Uncertainty in measuring volume. The uncertainty depends on the nature of the measuring device. Eight mL of liquid can be measured with less uncertainty in the 10-mL graduated cylinder than in the 100-mL graduated cylinder.

experiment or judge its precision. The three volume measurements referred to earlier could be reported as

8 ± 1 mL	(large graduated cylinder)
8.0 ± 0.1 mL	(small graduated cylinder)
8.00 ± 0.01 mL	(buret)

In this text, we will drop the ± notation and simply write

8 mL 8.0 mL 8.00 mL

There's a big difference between 8 mL and 8.00 mL, perhaps as much as half a milliliter.

When we do this, it is understood that there is an ***uncertainty of at least one unit in the last digit***—that is, 1 mL, 0.1 mL, 0.01 mL, respectively. This method of citing the degree of confidence in a measurement is often described in terms of **significant figures,** the meaningful digits obtained in a measurement. In 8.00 mL there are three significant figures; each of the three digits has experimental meaning. Similarly, there are two significant figures in 8.0 mL and one significant figure in 8 mL.

Frequently we need to know the number of significant figures in a measurement reported by someone else (Example 1.2).

EXAMPLE 1.2

Using different balances, three different students weigh the same object. They report the following masses:

(a) 1.611 g (b) 1.60 g (c) 0.001611 kg

How many significant figures does each value have?

STRATEGY

Assume each student reported the mass in such a way that the last number indicates the uncertainty associated with the measurement.

SOLUTION

(a) 1.611 g	4	
(b) 1.60 g	3	The zero after the decimal point is significant. It indicates that the object was weighed to the nearest 0.01 g.
(c) 0.001611 kg	4	The zeros at the left are not significant. They are only there because the mass was expressed in kilograms rather than grams. Note that 1.611 g and 0.001611 kg represent the same mass.

END POINT

If you express these masses in exponential notation as 1.611×10^{0} g, 1.60×10^{0} g, and 1.611×10^{-3} kg, the number of significant figures becomes obvious.

Sometimes the number of significant figures in a reported measurement is ambiguous. Suppose that a piece of metal is reported to weigh 500 g. You cannot be sure how many of these digits are meaningful. Perhaps the metal was

Unfortunately, the uncertainty here is uncertain.

weighed to the nearest gram (500 ± 1 g). If so, the 5 and the two zeros are significant; there are three significant figures. Then again, the metal might have been weighed only to the nearest 10 g (500 ± 10 g). In this case, only the 5 and one zero are known accurately; there are two significant figures. About all you can do in such cases is to wish the person who carried out the weighing had used exponential notation. The mass should have been reported as

$$5.00 \times 10^2 \text{ g} \qquad \text{(3 significant figures)}$$

or

$$5.0 \times 10^2 \text{ g} \qquad \text{(2 significant figures)}$$

or

$$5 \times 10^2 \text{ g} \qquad \text{(1 significant figure)}$$

In general, *any ambiguity concerning the number of significant figures in a measurement can be resolved by using exponential notation* (often referred to as "scientific notation"), discussed in Appendix 3.

The number of significant figures is the number of digits shown when a quantity is expressed in exponential notation.

Most measured quantities are not end results in themselves. Instead, they are used to calculate other quantities, often by multiplication or division. The precision of any such derived result is limited by that of the measurements on which it is based. ***When measured quantities are multiplied or divided, the number of significant figures in the result is the same as that in the quantity with the smallest number of significant figures.***

The rule is approximate, but sufficient for our purposes.

EXAMPLE 1.3

A patient is given an antibiotic intravenously. The rate of infusion is set so that the patient receives 1.15 mg of antibiotic per minute. How many milligrams of antibiotic are received after 35 minutes of infusion?

ANALYSIS

Information given:	rate of infusion (1.15 mg/min) time elapsed (35 minutes)
Asked for:	amount of antibiotic infused

STRATEGY

1. Substitute into the formula.

$$\text{rate} = \frac{\text{amount infused}}{\text{time}}$$

2. Recall the rules for significant figures.

SOLUTION

amount infused	amount infused = rate × time = (1.15 mg/min)(35 min) = 40.25 mg
significant figures	rate: 3; time: 2 The answer should have 2 significant figures.
amount infused	4.0×10^1 mg

The rules for "rounding off" a measurement, which were applied in Example 1.3, are as follows:

1. *If the digits to be discarded are less than – – 500 . . ., leave the last digit unchanged.* Masses of 23.315 g and 23.487 g both round off to 23 g if only two significant digits are required.

2. *If the digits to be discarded are greater than – – 500 . . ., add one to the last digit.* Masses of 23.692 g and 23.514 g round off to 24 g.
3. *If, perchance, the digits to be discarded are – – 500 . . . (or simply – – 5 by itself), round off so that the last digit is an even number.* Masses of 23.500 g and 24.5 g both round off to 24 g (two significant figures).

This way, you round up as often as you round down.

When measured quantities are added or subtracted, the uncertainty in the result is found in a quite different way than when they are multiplied and divided. It is determined by counting the number of decimal places, that is, the number of digits to the right of the decimal point for each measured quantity. ***When measured quantities are added or subtracted, the number of decimal places in the result is the same as that in the quantity with the greatest uncertainty and hence the smallest number of decimal places.***

In a multi-step calculation, round off only in the final step.

To illustrate this rule, suppose you want to find the total volume of a vanilla latté made up of 2 shots of espresso (1 shot = 46.1 mL), 301 mL of milk, and 2 tablespoons of vanilla syrup (1 tablespoon = 14.787 mL).

	Volume	Uncertainty	
Espresso coffee	92.2 mL	±0.1 mL	1 decimal place
Milk	301 mL	±1 mL	0 decimal place
Vanilla syrup	29.574 mL	±0.001 mL	3 decimal places
Total volume	423 mL		

Because there are no digits after the decimal point in the volume of milk, there are none in the total volume. Looking at it another way, we can say that the total volume, 423 mL, has an uncertainty of ±1 mL, as does the volume of milk, the quantity with the greatest uncertainty.

In applying the rules governing the use of significant figures, you should keep in mind that certain numbers involved in calculations are exact rather than approximate. To illustrate this situation, consider the equation relating Fahrenheit and Celsius temperatures:

$$t_{°F} = 1.8t_{°C} + 32°$$

The numbers 1.8 and 32 are exact. Hence they do not limit the number of significant figures in a temperature conversion; that limit is determined only by the precision of the thermometer used to measure temperature.

A different type of exact number arises in certain calculations. Suppose you are asked to determine the amount of heat evolved when *one kilogram* of coal burns. The implication is that because "one" is spelled out, *exactly* one kilogram of coal burns. The uncertainty in the answer should be independent of the amount of coal.

A number that is spelled out (one, two, . . .) does not affect the number of significant figures.

Conversion of Units

It is often necessary to convert a measurement expressed in one unit to another unit in the same system or to convert a unit in the English system to one in the metric system. To do this we follow what is known as a **conversion factor** approach or dimensional analysis. For example, to convert a volume of 536 cm^3 to liters, the relation

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

There are exactly 1000 cm^3 in exactly 1 L.

is used. Dividing both sides of this equation by 1000 cm^3 gives a quotient equal to 1:

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

The quotient 1 L/1000 cm^3, which is called a conversion factor, is multiplied by 536 cm^3. Because the conversion factor equals 1, this does not change the actual volume. However, it does accomplish the desired conversion of units. The cm^3 in the numerator and denominator cancel to give the desired unit: liters.

$$536\ \cancel{\text{cm}^3} \times \frac{1\ \text{L}}{1000\ \cancel{\text{cm}^3}} = 0.536\ \text{L}$$

To convert a volume in liters, say 1.28 L to cm^3, you must use a different form of the conversion factor. Use the units as a guide.

$$1.28\ \cancel{\text{L}} \times \frac{1000\ \text{cm}^3}{1\ \cancel{\text{L}}} = 1280\ \text{cm}^3 = 1.28 \times 10^3\ \text{cm}^3$$

Notice that a single relation (1 L = 1000 cm^3) gives two conversion factors:

$$\frac{1\ \text{L}}{1000\ \text{cm}^3} \quad \text{and} \quad \frac{1000\ \text{cm}^3}{1\ \text{L}}$$

Always check the units of your final answer. If you accidentally use the wrong form of the conversion factor, you will not get the desired unit. For example, if in your conversion of 1.28 L to cm^3 you used the conversion factor 1 L/1000 cm^3, you would get

$$1.28\ \text{L} \times \frac{1\ \text{L}}{1000\ \text{cm}^3} = 1.28 \times 10^{-3}\ \frac{\text{L}^2}{\text{cm}^3}$$

In general, when you make a conversion *choose the factor that cancels out the initial unit:*

$$\cancel{\text{initial unit}} \times \frac{\text{wanted unit}}{\cancel{\text{initial unit}}} = \text{wanted unit}$$

Conversions between English and metric units can be made using Table 1.3. We will call these "bridge conversions." They allow you to move from one system to another.

Table 1.3 Relations Between Length, Volume, and Mass Units

Metric		English		Metric-English	
Length					
1 km	= 10^3 m	1 ft	= 12 in	1 in	= 2.54 cm*
1 cm	= 10^{-2} m	1 yd	= 3 ft	1 m	= 39.37 in
1 mm	= 10^{-3} m	1 mi	= 5280 ft	1 mi	= 1.609 km
1 nm	= 10^{-9} m = 10 Å				
Volume					
1 m^3	= 10^6 cm^3 = 10^3 L	1 gal	= 4 qt = 8 pt	1 ft^3	= 28.32 L
1 cm^3	= 1 mL = 10^{-3} L	1 qt (U.S. liq)	= 57.75 in^3	1 L	= 1.057 qt (U.S. liq)
Mass					
1 kg	= 10^3 g	1 lb	= 16 oz	1 lb	= 453.6 g
1 mg	= 10^{-3} g	1 short ton	= 2000 lb	1 g	= 0.03527 oz
1 metric ton	= 10^3 kg			1 metric ton	= 1.102 short ton

*This conversion factor is exact; the inch is defined to be exactly 2.54 cm. The other factors listed in this column are approximate, quoted to four significant figures. Additional digits are available if needed for very accurate calculations. For example, the pound is defined to be 453.59237 g.

EXAMPLE 1.4

A red blood cell has a diameter of 7.5 μm (micrometers). What is the diameter of the cell in inches? (1 inch = 2.54 cm)

ANALYSIS

Information given:	cell diameter (7.5 μm) bridge conversion (1 in = 2.54 cm)
Information implied:	relation between micrometers and centimeters
Asked for:	7.5 μm in inches

STRATEGY

Follow the plan:	$\mu m \longrightarrow cm \longrightarrow inch$

SOLUTION

7.5 μm in inches	$7.5\ \mu m \times \frac{1 \times 10^{-6}\ m}{1\ \mu m} \times \frac{100\ cm}{1\ m} \times \frac{1\ in}{2.54\ cm} = \boxed{3.0 \times 10^{-4}\ in}$

Sometimes the required conversion has units raised to a power. To obtain the desired unit, you must remember to raise *both* the unit and the number to the desired power. Example 1.5 illustrates this point.

EXAMPLE 1.5

The beds in your dorm room have extra-long mattresses. These mattresses are 80 inches (2 significant figures) long and 39 inches wide. (Regular twin beds are 72 inches long.)

What is the area of the mattress top in m^2? (1 inch = 2.54 cm)

ANALYSIS

Information given:	mattress length (80 in) and width (39 in) bridge conversion (1 inch = 2.54 cm)
Information implied:	centimeter to meter conversion
Asked for:	area in m^2

STRATEGY

1. Recall equation for finding the area of a rectangle: area = length × width
2. Follow the plan: $in^2 \longrightarrow cm^2 \longrightarrow m^2$

SOLUTION

area in in^2	$80\ in \times 39\ in = 3.12 \times 10^3\ in^2$ (We will round off to correct significant figures at the end.)
area in m^2	$3.12 \times 10^3\ in^2 \times \frac{(2.54)^2\ cm^2}{(1)^2\ in^2} \times \frac{(1)^2\ m^2}{(100)^2\ cm^2} = \boxed{2.0\ m^2}$

END POINT

There are 36 inches in one yard, so the dimensions of the mattresses are approximately 1 yd wide and 2 yd long or 2 yd^2. A meter is almost equivalent to a yard (see Table 1.3) so the calculated answer is in the same ball park.

CHEMISTRY THE HUMAN SIDE

The discussion in Section 1-2 emphasizes the importance of making precise numerical measurements. Chemistry was not always so quantitative. The following recipe for finding the philosopher's stone was recorded more than 300 years ago.

> Take all the mineral salts there are, also all salts of animal and vegetable origin. Add all the metals and minerals, omitting none. Take two parts of the salts and grate in one part of the metals and minerals. Melt this in a crucible, forming a mass that reflects the essence of the world in all its colors. Pulverize this and pour vinegar over it. Pour off the red liquid into English wine bottles, filling them half-full. Seal them with the bladder of an ox (*not* that of a pig). Punch a hole in the top with a coarse needle. Put the bottles in hot sand for three months. Vapor will escape through the hole in the top, leaving a red powder. . . .

One man more than any other transformed chemistry from an art to a science. Antoine Lavoisier was born in Paris; he died on the guillotine during the French Revolution. (Lavoisier was executed because he was a tax collector; chemistry had nothing to do with it.) Above all else, Lavoisier understood the importance of carefully controlled, quantitative experiments. These were described in his book *Elements of Chemistry*. Published in 1789, it is illustrated with diagrams by his wife. (E. I. DuPont (1772–1834) was a student of Lavoisier.)

Science Photo Library/Custom Medical Stock

Antoine Lavoisier
(1743–1794)

The results of one of Lavoisier's quantitative experiments are shown in Table A; the data are taken directly from Lavoisier. If you add up the masses of reactants and products (expressed in arbitrary units), you find them to be the same, 510. As Lavoisier put it, "In all of the operations of men and nature, nothing is created. An equal quantity of matter exists before and after the experiment."

This was the first clear statement of the law of conservation of mass (Chapter 2), which was the cornerstone for the growth of chemistry in the nineteenth century. Again, to quote Lavoisier, "it is on this principle that the whole art of making experiments is founded."

Table A Quantitative Experiment on the Fermentation of Wine (Lavoisier)

Reactants	Mass (Relative)	Products	Mass (Relative)
Water	400	Carbon dioxide	35
Sugar	100	Alcohol	58
Yeast	10	Acetic acid	3
		Water	409
		Sugar (unreacted)	4
		Yeast (unreacted)	1

1-3 Properties of Substances

Every pure substance has its own unique set of properties that serve to distinguish it from all other substances. A chemist most often identifies an unknown substance by measuring its properties and comparing them with the properties recorded in the chemical literature for known substances.

The properties used to identify a substance must be **intensive properties;** that is, they must be independent of amount. The fact that a sample weighs 4.02 g or has a volume of 229 mL tells us nothing about its identity; mass and volume are **extensive properties**; that is, they depend on amount. Beyond that, substances may be identified on the basis of their

- **chemical properties,** observed when the substance takes part in a chemical reaction, a change that converts it to a new substance. For example, the fact that mercury(II) oxide decomposes to mercury and oxygen on heating to 600°C can be used to identify it. Again, the chemical inertness of helium helps to distinguish it from other, more reactive gases, such as hydrogen and oxygen.

- **physical properties,** observed without changing the chemical identity of a substance. Two such properties particularly useful for identifying a substance are
  - **melting point,** the temperature at which a substance changes from the solid to the liquid state.
  - **boiling point,** the temperature at which bubbles filled with vapor form within a liquid. If a substance melts at 0°C and boils at 100°C, we are inclined to suspect that it might just be water.

Taste is a physical property, but it is never measured in the lab.

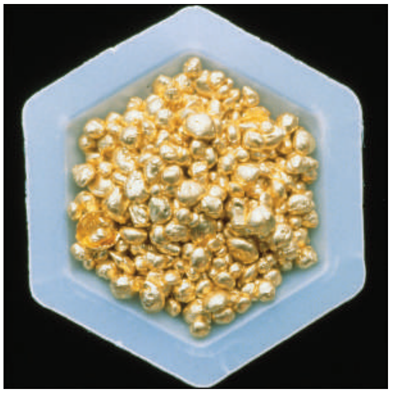

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Figure 1.12 Properties of gold. The color of gold is an *intensive* property. The quantity of gold in a sample is an *extensive* property. The fact that gold can be stored in the air without undergoing any chemical reaction with oxygen in the air is a *chemical* property. The temperature at which gold melts (1063°C) is a *physical* property.

Figure 1.12 shows gold and describes its different properties. In the remainder of this section we will consider a few other physical properties that can be measured without changing the identity of a substance.

Density

The **density** of a substance is the ratio of mass to volume:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad d = \frac{\text{mass}}{V} \tag{1.3}$$

Note that even though mass and volume are extensive properties, the ratio of mass to volume is intensive. Samples of copper weighing 1.00 g, 10.5 g, 264 g, . . . all have the same density, 8.94 g/mL at 25°C.

For liquids and gases, density can be found in a straightforward way by measuring independently the mass (using a scale) and the volume (using a pipet or graduated cylinder) of a sample. (Example 1.6 illustrates the process.)

EXAMPLE 1.6

Glycerol is a viscous liquid used by both the pharmaceutical and food industries as a sweetener, thickener, and stabilizer. To determine its density, a student delivers a 15.0 mL sample by pipet into a flask with a mass of 28.45 g. The mass of the flask and glycerol sample is 47.37 g. What is the density of glycerol?

ANALYSIS

Information given:	mass of empty flask (28.45 g) mass of flask + sample (47.37 g) volume of sample (15.0 mL)
Asked for:	density of the sample

STRATEGY

1. Find the mass of the sample by difference.

$$\text{mass of sample} = (\text{mass of flask} + \text{sample}) - (\text{mass of flask})$$

2. Recall the formula for density.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

SOLUTION

1. mass of sample	$\text{mass of sample} = (\text{mass of flask} + \text{sample}) - (\text{mass of flask}) = 47.37\text{ g} - 28.45\text{ g} = 18.92\text{ g}$
2. density	$d = \frac{\text{mass}}{V} = \frac{18.92\text{ g}}{15.0\text{ mL}} = \boxed{1.26\text{ g/mL}}$

END POINT

The density is expressed in three significant figures because the volume is measured only to 3 significant figures.

For solids, the mass of the sample can be obtained directly by weighing it. The volume of a regular solid (a cube, for example) can be calculated using the given dimensions of the sample. The volume of an irregular solid, like a rock, is obtained by displacement (Example 1.7).

EXAMPLE 1.7

Consider two samples of palladium (Pd), an element used in automobile catalytic converters. Sample A is a cylindrical bar with a mass of 97.36 g. The bar is 10.7 cm high and has a radius of 4.91 mm. Sample B is an irregular solid with a mass of 49.20 g. A graduated cylinder has 10.00 mL of water. When sample B is added to the graduated cylinder, the volume of the water and the solid is 14.09 mL. Calculate the density of each sample.

SAMPLE A:

ANALYSIS

Information given:	mass (97.36 g), radius, r (4.91 mm), height, h (10.7 cm)
Asked for:	density of Pd

STRATEGY

1. Recall the formula to obtain the volume of a cylinder.

$$V = \pi r^2 h$$

2. Substitute into the definition of density.

$$d = \frac{\text{mass}}{V}$$

SOLUTION

V	$V = \pi r^2 h = \pi\left(4.91\text{ mm} \times \frac{1\text{ cm}}{10\text{ mm}}\right)^2 \times 10.7\text{ cm} = 8.10\text{ cm}^3$
d	$d = \frac{\text{mass}}{V} = \frac{97.36\text{ g}}{8.10\text{ cm}^3} = \boxed{12.0\text{ g/cm}^3}$

SAMPLE B:

ANALYSIS

Information given:	mass: (49.20 g) volume of water before Pd addition: (10.00 mL) volume of water and Pd: (14.09 mL)
Asked for:	density of Pd

STRATEGY

1. The volume of the Pd is the difference between the volume of the Pd and water and the volume of the water alone.
2. Substitute into the definition of density.

$$d = \frac{\text{mass}}{V}$$

SOLUTION

V	$V = V_{H_2O + Pd} - V_{H_2O} = 14.09\text{ mL} - 10.00\text{ mL} = 4.09\text{ mL}$
d	$d = \frac{\text{mass}}{V} = \frac{49.02\text{ g}}{4.09\text{ mL}} = \boxed{12.0\text{ g/mL}}$

END POINT

The units for density are g/cm^3 and g/mL. Since 1 cm^3 = 1 mL, these can be used interchangeably.

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Figure 1.13 Density. The wood block has a lower density than water and floats. The ring has a higher density than water and sinks.

Figure 1.13 shows the relation between substances with a density greater than that of water (1.0 g/mL) and those with a density less than that of water.

In a practical sense, density can be treated as a conversion factor to relate mass and volume. Knowing that mercury has a density of 13.6 g/mL, we can calculate the mass of 2.6 mL of mercury:

$$2.6\ \cancel{\text{mL}} \times 13.6\ \frac{\text{g}}{\cancel{\text{mL}}} = 35\ \text{g}$$

or the volume occupied by one kilogram of mercury:

$$1.000\ \cancel{\text{kg}} \times \frac{10^3\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \text{mL}}{13.6\ \cancel{\text{g}}} = 73.5\ \text{mL}$$

Solubility

Solubility is referred to as the process by which a solute dissolves in a solvent, and is ordinarily a physical rather than a chemical change. The extent to which it dissolves can be expressed in various ways. A common method is to state the number of grams of the substance that dissolves in 100 g of solvent at a given temperature.

EXAMPLE 1.8 GRADED

Sucrose is the chemical name for the sugar we consume. Its solubility at 20°C is 204 g/100 g water, and at 100°C is 487 g/100 g water. A solution is prepared by mixing 139 g of sugar in 33.0 g of water at 100°C.

a What is the minimum amount of water required to dissolve the sugar at 100°C?

ANALYSIS

Information given:	sucrose solubility at 100°C (487 g/100 g water) composition of solution: sucrose (139 g), water (33.0 g)
Asked for:	minimum amount of H_2O to dissolve 139 g sucrose at 100°C

STRATEGY

Relate the mass H_2O required to the mass sucrose to be dissolved at 100°C by using the solubility at 100°C as a conversion factor.

SOLUTION

mass H_2O required

$$139\ \text{g sucrose} \times \frac{100\ \text{g}\ H_2O}{487\ \text{g sucrose}} = 28.5\ \text{g}\ H_2O$$

b What is the maximum amount of sugar that can be dissolved in the water at 100°C?

ANALYSIS

Information given:	sucrose solubility at 100°C (487 g/100 g water) composition of solution: sucrose (139 g), water (33.0 g)
Asked for:	maximum amount of sucrose that can be dissolved in 33.0 g H_2O at 100°C

STRATEGY

Relate the mass of H_2O required to the mass of sucrose to be dissolved at 100°C by using the solubility at 100°C as a conversion factor.

SOLUTION

mass sucrose

$$33.0\ \text{g}\ H_2O \times \frac{487\ \text{g sucrose}}{100\ \text{g}\ H_2O} = 161\ \text{g sucrose}$$

continued

C The solution is cooled to 20°C. How much sugar (if any) will crystallize out?	
ANALYSIS	
Information given:	sucrose solubility at 20°C (204 g/100 g water) composition of solution: sucrose (139 g), water (33.0 g)
Asked for:	mass of sucrose in the solution that will not dissolve at 20°C
STRATEGY	
1. The question really is: How much sucrose will dissolve in 33.0 g water at 20°C? Relate the mass of sucrose that can be dissolved by 33.0 g H_2O at 20°C by using the solubility at 20°C as a conversion factor. **2.** Take the difference between the calculated amount that can be dissolved and the amount of sucrose that was in solution at 100°C.	
SOLUTION	
1. mass sucrose	$33.0 \text{ g } H_2O \times \dfrac{204 \text{ g sucrose}}{100 \text{ g } H_2O} = 67.3 \text{ g sucrose will dissolve at 20°C.}$
2. undissolved sucrose	139 g in solution − 67.3 g can be dissolved = 72 g undissolved

Figure 1.14 shows the solubility of sugar in water as a function of temperature. Alternatively, we can say that it gives the concentration of sugar in a **saturated solution** at various temperatures. For example, at 20°C, we could say that "the solubility of sugar is 204 g/100 g water" or that "a saturated solution of sugar contains 204 g/100 g water."

The "100 g water" in the solubility expression is an exact quantity.

At any point in the area below the curve in Figure 1.14, we are dealing with an **unsaturated solution**. Consider, for example, point A (150 g sugar per 100 g water at 20°C). This solution is unsaturated; if we add more sugar, another 54 g will dissolve to give a saturated solution (204 g sugar per 100 g water at 20°C).

At any point in the area above the curve, the sugar solution is **supersaturated solution.** This is the case at point B (300 g sugar per 100 g water at 20°C). Such a solution could be formed by carefully cooling a saturated solution at 60°C to 20°C, where a saturated solution contains 204 g sugar per 100 g water. The excess sugar stays in solution until a small seed crystal of sugar is added, whereupon crystallization quickly takes place. At that point the excess sugar

$$300 \text{ g} - 204 \text{ g} = 96 \text{ g}$$

comes out of solution.

When the temperature changes, the amount of solute in solution changes, but the mass of water stays the same.

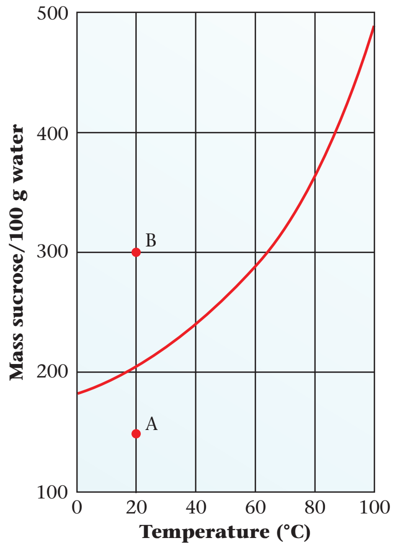

Figure 1.14 Solubility of table sugar (sucrose). The solubility of sugar, $C_{12}H_{22}O_{11}$, in water increases exponentially with temperature.

CHEMISTRY **BEYOND** THE CLASSROOM

Arsenic

An element everyone has heard about but almost no one has ever seen is arsenic, symbol As. It is a gray solid with some metallic properties, melts at 816°C, and has a density of 5.78 g/mL. Among the elements, arsenic ranks 51st in abundance. It is about as common as tin or beryllium. Two brightly colored sulfides of arsenic, realgar and orpiment (Figure A), were known to the ancients. The element is believed to have been isolated for the first time by Albertus Magnus in the thirteenth century. He heated orpiment with soap. The alchemists gave what they thought to be arsenic (it really was an oxide of arsenic) its own symbol (Figure B) and suggested that women rub it on their faces to whiten their complexion.

The principal use of elemental arsenic is in its alloys with lead. The "lead" storage battery contains a trace of arsenic along with 3% antimony. Lead shot, which are formed by allowing drops of molten lead to fall through the air, contains from 0.5 to 2.0% arsenic. The presence of arsenic raises the surface tension of the liquid and hence makes the shot more spherical.

In the early years of the twentieth century, several thousand organic compounds were synthesized and tested for medicinal use, mainly in the treatment of syphilis. One of these compounds, salvarsan, was found to be very effective. Arsenic compounds fell out of use in the mid-twentieth century because of the unacceptable side effects that occurred at the dosages that were thought to be necessary. In the 1970s, Chinese medicine tried a highly purified oxide of arsenic in a low-dose regimen. It was shown to be effective in the treatment of some leukemias. Western medicine has confirmed these results. Molecular studies and clinical trials are ongoing and suggest that arsenic oxides show great promise in the treatment of malignant disease.

Industry and farming have used arsenic compounds. However, because of its great toxicity and ability to leach into wells and streams, its is no longer produced in the United States, but it is still imported from other countries. Until the 1940s, arsenic compounds were used as agricultural pesticides. Today, most uses of arsenic in farming are banned in the United States, and its use as a preservative in pressure-treated wood has been greatly reduced.

The "arsenic poison" referred to in crime dramas is actually an oxide of arsenic rather than the element itself. Less than 0.1 g of this white, slightly soluble powder can be fatal. The classic symptoms of arsenic poisoning involve various unpleasant gastrointestinal disturbances, severe abdominal pain, and burning of the mouth and throat.

Gary Cook, Inc/Visuals Unlimited/Encyclopedia/Corbis

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Figure A **Realgar and orpiment.**

In the modern forensic laboratory, arsenic is detected by analysis of hair samples. A single strand of hair is sufficient to establish the presence or absence of the element. The technique most commonly used is neutron activation analysis, described in Chapter 18. If the concentration is found to be greater than about 0.0003%, poisoning is indicated.

This technique was applied in the early 1960s to a lock of hair taken from Napoleon Bonaparte (1769–1821) on St. Helena. Arsenic levels of up to 50 times normal suggested he may have been a victim of poisoning, perhaps on orders from the French royal family. More recently (1991), U.S. President Zachary Taylor (1785–1850) was exhumed on the unlikely hypothesis that he had been poisoned by Southern sympathizers concerned about his opposition to the extension of slavery. The results indicated normal arsenic levels. Apparently, he died of cholera, brought on by an overindulgence in overripe and unwashed fruit.

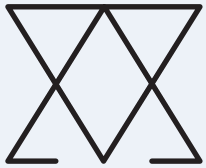


Figure B **Alchemist symbol for arsenic.**

Marna G. Clarke

The crystallization of excess solute is a common problem in the preparation of candies and in the storage of jam and honey. From these supersaturated solutions, sugar separates either as tiny crystals, causing the "graininess" in fudge, or as large crystals, which often appear in honey kept for a long time (Figure 1.15).

Figure 1.15 **Rock candy.** The candy is formed by crystallization of sugar from a saturated solution that is cooled slowly.

Key Concepts

1. Convert between °F, °C, and K.
   (**Example 1.1; Problems 13–16**)
2. Determine the number of significant figures in a measured quantity.
   (**Example 1.2; Problems 17, 18, 25, 26**)
3. Determine the number of significant figures in a calculated quantity.
   (**Example 1.3; Problems 27–30**)
4. Use conversion factors to change the units of a measured quantity.
   (**Example 1.4; Problems 31–44, 60**)
5. Relate density to mass and volume.
   (**Example 1.5; Problems 45–52, 61, 63**)
6. Given its solubility, relate mass of solute to that of solvent.
   (**Example 1.6; Problems 53–56, 57**)

Key Equations

Fahrenheit temperature $t_{°F} = 1.8\, t_{°C} + 32°$

Kelvin temperature $T_K = t_{°C} + 273.15$

Density $\dfrac{\text{mass}}{V}$

Key Terms

centi
compound
conversion factor
chemical properties
density
element
extensive properties
intensive properties
kilo
milli
mixture
—heterogeneous
—homogeneous
nano
physical properties
significant figures
solution
—saturated solution
—supersaturated solution
—unsaturated solution

Summary Problem

Potassium permanganate, a purple crystal, can be used as a mild antiseptic to prevent infection from cuts and scrapes. It is made up of three atoms: potassium, oxygen, and manganese. Consider the following information about potassium permanganate:

- density at 20°C: 2.703 g/cm^3
- melting point: 2.40×10^2°C
- solubility at 20°C: 6.38 g/100 g water
- solubility at 60°C: 25 g/100 g water

(a) What are the symbols of the three elements in potassium permanganate?

(b) List the physical properties of potassium permanganate described above.

(c) What is the mass of a sample of potassium permanganate with a volume of 48.7 cm^3?

(d) Express the density in pounds per cubic foot.

(e) Express the melting point of potassium permanganate in °F and K.

(f) How many grams of potassium permanganate can be dissolved in 38.5 g of water at 20°C?

(g) An aqueous solution of potassium permanganate is prepared by dissolving 15.0 g of potassium permanganate in 65.0 g of water at 60°C. The solution is carefully cooled to 20°C. A homogeneous solution is obtained. Is the solution at 60°C saturated, unsaturated, or supersaturated? What about the solution at 20°C?

(h) A solution is prepared by dissolving 10.0 g of potassium permanganate in 55.0 g of water at 60°C. Will all the solute dissolve? How many grams of solute will crystallize out of solution if it is rapidly cooled to 20°C?

Express your answers to the correct number of significant figures; use the conversion factor method throughout.

Answers

(a) **K, Mn, O**
(b) **density, melting point, solubility, color**
(c) **132 g**
(d) **168.7 lb/ft^3**
(e) **464°F; 513 K**
(f) **2.46 g**
(g) **at 60°C: unsaturated; at 20°C: supersaturated**
(h) **yes; 6.5 g**

Questions and Problems

Even-numbered questions and Challenge Problems have answers in Appendix 5 and fully worked solutions in the *Student Solutions Manual*.

The questions and problems listed here are typical of those at the end of each chapter. Some are conceptual. Most require calculations, writing equations, or other quantitative work. The headings identify the primary section and topic of each set of questions or problems, such as "Symbols and Formulas" or "Significant Figures." Those in the "Unclassified" category may involve more than one concept, including, perhaps, topics from a preceding chapter. "Challenge Problems," listed at the end of the set, require extra skill and/or effort. The "Classified" questions and problems (Problems 1–56 in this set) occur in matched pairs, one below the other, and illustrate the same concept. For example, Questions 1 and 2 are nearly identical in nature; the same is true of Questions 3 and 4, and so on.

1-1 Matter and Its Classifications

Types of Matter

1. Classify each of the following as element, compound, or mixture.
(a) air (b) iron
(c) soy sauce (d) table salt

2. Classify each of the following as element, compound, or mixture.
(a) gold
(b) milk
(c) sugar
(d) vinaigrette dressing with herbs

3. Classify the following as solution or heterogeneous mixture.
(a) normal urine
(b) gasoline
(c) batter for chocolate chip cookies

4. Classify the following as solution or heterogeneous mixture.
(a) iron ore
(b) chicken noodle soup
(c) tears

5. How would you separate into its components
(a) a homogeneous solution of table salt and water?
(b) sand from gasoline?
(c) carbon dioxide gas from butane gas?

6. How would you separate into its different components
(a) a mixture of the volatile gases propane, butane, and isopropane?
(b) a solution of rubbing alcohol made up of isopropyl alcohol and water?

7. Write the symbol for the following elements.
(a) titanium (b) phosphorus
(c) potassium (d) magnesium

8. Write the symbol for the following elements.
(a) copper (b) carbon
(c) bromine (d) aluminum

9. Write the name of the element represented by the following symbols.
(a) Hg (b) Si (c) Na (d) I

10. Write the name of the element represented by the following symbols:
(a) Mn (b) Au (c) Pb (d) Ag

1-2 Measurements

Measurements

11. What instrument would you use to determine
(a) the mass of a head of lettuce?
(b) whether your refrigerator is cooling water to 10°C?
(c) the volume of a glass of orange juice?

12. What instrument would you use to
(a) measure the length of your bed?
(b) measure the amount of acid delivered by a beaker?
(c) determine whether you have a fever?

13. A glass of lukewarm milk is suggested for people who cannot sleep. Milk at 52°C can be characterized as lukewarm. What is the temperature of lukewarm milk in °F? In K?

14. A recipe for apple pie calls for a preheated 350°F (three significant figures) oven. Express this temperature setting in °C and in K.

15. Gallium is one of the few metals that can melt at room temperature. Its melting point is 29.76°C. If you leave solid gallium in your car on an early summer morning when the temperature is 75.0°F, what physical state is the gallium in when you return to your car and the interior car temperature is 85.0°F?

16. Computers are not supposed to be in very warm rooms. The highest termperature tolerated for maximum performance is 308 K. Express this temperature in °C and °F.

Significant Figures

17. How many significant figures are in each of the following?
(a) 0.890 cm (b) 210°C
(c) 2.189×10^6 nm (d) 2.54 cm = 1 in
(e) 140.00 g

18. How many significant figures are there in each of the following?
(a) 0.136 m (b) 0.0001050 g
(c) 2.700×10^3 nm (d) 6×10^{-4} L
(e) 56003 cm^3

19. Round off the following quantities to the indicated number of significant figures.
(a) 7.4855 g (three significant figures)
(b) 298.693 cm (five significant figures)
(c) 11.698 lb (one significant figure)
(d) 12.05 oz (three significant figures)

20. Round off the following quantities to the indicated number of significant figures.
(a) 17.2509 cm (4 significant figures)
(b) 168.51 lb (3 significant figures)
(c) 500.22°C (3 significant figures)
(d) 198.500 oz (3 significant figures)

21. Express the following in scientific notation:
(**a**) 132.5 cm
(**b**) 0.000883 km
(**c**) 6,432,000,000 nm

22. Express the following measurements in scientific notation.
(**a**) 4020.6 mL (**b**) 1.006 g (**c**) 100.1°C

23. Which of the following statements use only exact numbers?
(**a**) The little boy weighs 74.3 lbs.
(**b**) The boiling point of ethanol is 78°C.
(**c**) The mailman delivered 3 letters, 2 postcards and one magazine.

24. Which of the following statements use only exact numbers?
(**a**) The temperature in our dorm room is kept at 72°F.
(**b**) I bought 6 eggs, 2 cookies, and 5 tomatoes at the farmers' market.
(**c**) There are 1×10^9 nanometers in 1 meter.

25. A basketball game at the University of Connecticut's Gampel Pavilion attracted 10,000 people. The building's interior floor space has an area of 1.71×10^5 ft^2. Tickets to the game sold for $22.00. Senior citizens were given a 20% discount. How many significant figures are there in each quantity? (Your answer may include the words *ambiguous* and *exact*.)

26. A listing of a house for sale states that there are 5 bedrooms, 4000 ft^2 of living area, and a living room with dimensions 17×18.5 ft. How many significant figures are there in each quantity? (Your answer may include the words *ambiguous* and *exact*.)

27. Calculate the following to the correct number of significant figures. Assume that all these numbers are measurements.
(**a**) $x = 17.2 + 65.18 - 2.4$
(**b**) $x = \dfrac{13.0217}{17.10}$
(**c**) $x = (0.0061020)(2.0092)(1200.00)$
(**d**) $x = 0.0034 + \dfrac{\sqrt{(0.0034)^2 + 4(1.000)(6.3 \times 10^{-4})}}{(2)(1.000)}$
(**e**) $x = \dfrac{(2.998 \times 10^8)(3.1 \times 10^{-7})}{6.022 \times 10^{23}}$

28. Perform the indicated calculations. Write your answers with the correct number of significant figures.
(**a**) $\dfrac{13.01\ \text{g}}{1.23\ \text{L}}$
(**b**) 17.2 lb + 65.18 lb − 2.3 lb
(**c**) $(0.003240\ \text{cm})(1.25\ \text{cm})(1.0 \times 10^2\ \text{cm})$
(**d**) $\dfrac{0.024 + \sqrt{(0.024)^2 + 4(6.3 \times 10^{-3})}}{2}$
Where 4 and 2 are exact numbers, and all other quantities are measurements.
(**e**) $\dfrac{(4.63 \times 10^4)(1.8 \times 10^{-6})}{(6.022 \times 10^{23})}$
Assume all quantities are measurements.

29. The volume of a sphere is $4\pi r^3/3$, where r is the radius. One student measured the radius to be 4.30 cm. Another measured the radius to be 4.33 cm. What is the difference in volume between the two measurements?

30. The volume of a square pyramid is (1/3)Bh where B is the area of the square base and h is the height. One student's measurements of a pyramid are B = 1.33 m^2, h = 2.79 m. Another student's measurements of the same pyramid are B = 1.31 m^2, h = 2.81 m. What is the difference in the volumes of the pyramid calculated by the two students?

Conversion Factors

31. Write the appropriate symbol in the blank (>, <, or =).
(**a**) 303 m ________ 303×10^3 km
(**b**) 500 g ________ 0.500 kg
(**c**) 1.50 cm^3 ________ 1.50×10^3 nm^3

32. Write the appropriate symbol in the blank (>, <, or =).
(**a**) 37.12 g ________ 0.3712 kg
(**b**) 28 m^3 ________ 28×10^2 cm^3
(**c**) 525 mm ________ 525×10^6 nm

33. Convert 22.3 mL to
(**a**) liters (**b**) in^3 (**c**) quarts

34. Convert 0.2156 L to
(**a**) mL
(**b**) in^3
(**c**) quarts

35. The height of a horse is usually measured in hands. One hand is exactly 1/3 ft.
(**a**) How tall (in feet) is a horse of 19.2 hands?
(**b**) How tall (in meters) is a horse of 17.8 hands?
(**c**) A horse of 20.5 hands is to be transported in a trailer. The roof of the trailer needs to provide 3.0 ft of vertical clearance. What is the minimum height of the trailer in feet?

36. At sea, distances are measured in nautical miles and speeds are expressed in knots.

1 nautical mile = 6076.12 ft
1 knot = 1 nautical mi/h (exactly)

(**a**) How many miles are in one nautical mile?
(**b**) How many meters are in one nautical mile?
(**c**) A ship is traveling at a rate of 22 knots. Express the ship's speed in miles per hour.

37. The unit of land measure in the English system is the acre, while that in the metric system is the hectare. An acre is 4.356×10^4 ft^2. A hectare is ten thousand square meters. A town requires a minimum area of 2.0 acres of land for a single-family dwelling. How many hectares are required?

38. A gasoline station in Manila, Philippines, charges 38.46 pesos per liter of unleaded gasoline at a time when one U.S. dollar (USD) buys 47.15 pesos (PHP). The car you are driving has a gas tank with a capacity of 14 U.S. gallons and gets 24 miles per gallon.
(**a**) What is the cost of unleaded gasoline in Manila in USD per gallon?
(**b**) How much would a tankful of unleaded gasoline for your car cost in USD?
(**c**) Suppose that you have only PHP 1255 (a day's wage for an elementary school teacher) and the car's tank is

almost empty. How many miles can you expect to drive if you spend all your money on gasoline?

39. A lap in most tracks in the United States is 0.25 mi (English lap). In most countries that use the metric system, a metric lap is 0.40 km. A champion marathon runner covers a mile in about 5.0 min. How many minutes will the runner take to run an English lap? A metric lap?

40. Cholesterol in blood is measured in milligrams of cholesterol per deciliter of blood. If the unit of measurement were changed to grams of cholesterol per milliliter of blood, what would a cholesterol reading of 185 mg/dL translate to?

41. Some states have reduced the legal limit for alcohol sobriety from 0.10% to 0.080% alcohol by volume in blood plasma.

(**a**) How many milliliters of alcohol are in 3.0 qt of blood plasma at the lower legal limit?

(**b**) How many milliliters of alcohol are in 3.0 qt of blood plasma at the higher legal limit?

(**c**) How much less alcohol is in 3.0 qt of blood plasma with the reduced sobriety level?

42. The area of the 48 contiguous states is 3.02×10^6 mi^2. Assume that these states are completely flat (no mountains and no valleys). What volume, in liters, would cover these states with a rainfall of two inches?

43. In Europe, nutritional information is given in kilojoules (kJ) instead of nutritional calories (1 nutritional calorie = 1 kcal). A packet of soup has the following nutritional information:

$$250 \text{ mL of soup} = 235 \text{ kJ}$$

How would that same packet be labeled in the United States if the information has to be given in nutritional calories per cup? (There are 4.18 joules in one calorie and 2 cups to a pint.)

44. In the old pharmaceutical system of measurements, masses were expressed in grains. There are 5.760×10^3 grains in 1 lb. An old bottle of aspirin lists 5.000 grains of active ingredient per tablet. How many milligrams of active ingredient are there in the same tablet?

1-3 Properties of Substances

Physical and Chemical Properties

45. The cup is a measure of volume widely used in cookbooks. One cup is equivalent to 225 mL. What is the density of clover honey (in grams per milliliter) if three quarters of a cup has a mass of 252 g?

46. The egg whites from four "large" eggs occupy a volume of 112 mL. The mass of the egg whites from these four eggs is 1.20×10^2 g. What is the average density of an egg white from one "large" egg?

47. A metal slug weighing 25.17 g is added to a flask with a volume of 59.7 mL. It is found that 43.7 g of methanol (d = 0.791 g/mL) must be added to the metal to fill the flask. What is the density of the metal?

48. A solid with an irregular shape and a mass of 11.33 g is added to a graduated cylinder filled with water (d = 1.00 g/mL) to the 35.0-mL mark. After the solid sinks to the bottom, the water level is read to be at the 42.3-mL mark. What is the density of the solid?

49. A waterbed filled with water has the dimensions 8.0 ft × 7.0 ft × 0.75 ft. Taking the density of water to be 1.00 g/cm^3, how many kilograms of water are required to fill the waterbed?

50. Wire is often sold in pound spools according to the wire gauge number. That number refers to the diameter of the wire. How many meters are in a 10-lb spool of 12-gauge aluminum wire? A 12-gauge wire has a diameter of 0.0808 in. Aluminum has a density of 2.70 g/cm^3. ($V = \pi r^2 \ell$)

51. Air is 21% oxygen by volume. Oxygen has a density of 1.31 g/L. What is the volume, in liters, of a room that holds enough air to contain 55 kg of oxygen?

52. The unit for density found in many density tables is kg/m^3. At a certain temperature, the gasoline you pump into your car's gas tank has a density of 732.22 kg/m^3. If your tank has a capacity of 14.0 gallons, how many grams of gasoline are in your tank when it is full? How many pounds?

53. Magnesium sulfate ($MgSO_4$) has a solubility of 38.9 g/100 g H_2O at 30°C. A solution is prepared by adding 9.50 g of $MgSO_4$ to 25.0 g of water at 40°C. A homogeneous mixture is obtained. Is the solution saturated, unsaturated, or supersaturated? One gram of magnesium sulfate is added to the solution cooled to 30°C. Would you expect some of the $MgSO_4$ to precipitate? If so, how much? If not, how much more $MgSO_4$ can be added before precipitation takes place?

54. Potassium sulfate has a solubility of 15 g/100 g water at 40°C. A solution is prepared by adding 39.0 g of potassium sulfate to 225 g of water, carefully heating the solution, and cooling it to 40°C. A homogeneous solution is obtained. Is this solution saturated, unsaturated, or supersaturated? The beaker is shaken, and precipitation occurs. How many grams of potassium sulfate would you expect to crystallize out?

55. Sodium bicarbonate (baking soda) is commonly used to absorb odor. Its solubility is 9.6 g/100 g H_2O at 30°C and 16 g/100 g H_2O at 60°C. At 60°C, 9.2 g of baking soda are added to 46 g of water.

(**a**) Is the resulting mixture homogeneous at 60°C? If not, how many grams of baking soda are undissolved?

(**b**) The mixture is cooled to 30°C. How many more grams of water are needed to make a saturated solution?

56. Magnesium chloride is an important coagulant used in the preparation of tofu from soy milk. Its solubility in water at 20°C is 54.6 g/100 g. At 80°C, its solubility is 66.1 g/100 g. A mixture is made up of 16.2 g of magnesium chloride and 38.2 g of water at 20°C.

(**a**) Is the mixture homogeneous? If it is, how many more grams of magnesium chloride are required to make a saturated solution? If the mixture is not homogeneous, how many grams of magnesium chloride are undissolved?

(**b**) How many more grams of magnesium chloride are needed to make a saturated solution at 80°C?

Unclassified

57. The solubility of lead nitrate at 100°C is 140.0 g/100 g water. A solution at 100°C consists of 57.0 g of lead nitrate in 64.0 g of water. When the solution is cooled to 10°C, 25.0 g of lead nitrate crystallize out. What is the solubility of lead nitrate in g/100 g water at 10°C?

58. Radiation exposure to human beings is usually given in rems (radiation equivalent for man). In SI units, the exposure is measured in sieverts (Sv). One rem equals 0.0100 Sv. At one time, the exposure due to the nuclear reactors in Japan was measured to be 8217 mSv/h. How many rems would a person exposed to the radiation for 35 min have absorbed? If one mammogram gives off 0.30 rems, how many mammograms would that exposure be equivalent to?

59. The following data refer to the element phosphorus. Classify each as a physical or a chemical property.

(a) It exists in several forms, for example, white, black, and red phosphorus.
(b) It is a solid at 25°C and 1 atm.
(c) It is insoluble in water.
(d) It burns in chlorine to form phosphorus trichloride.

60. A supersaturated sugar solution (650.0 g sugar in 150.0 g water) is prepared to make "rock candy." The solution is kept at 25°C, where the solubility of sugar is 220.0 g/100 g water. A string is dipped into the solution, and in time, large crystals of sugar (rock candy) form. How many grams of solid are obtained when the solution just reaches saturation?

61. The density of wind-packed snow is estimated to be 0.35 g/cm^3. A flat roof that is 35 by 43 feet has 28 inches of snow on it. How many pounds of snow are on the roof?

62. The dimensions of aluminum foil in a box for sale in supermarkets are 66⅔ yards by 12 inches. The mass of the foil is 0.83 kg. If its density is 2.70 g/cm^3, then what is the thickness of the foil in inches?

63. The Kohinoor Diamond (d = 3.51 g/cm^3) is 108 carats. If one carat has a mass of 2.00×10^2 mg, what is the mass of the Kohinoor Diamond in pounds? What is the volume of the diamond in cubic inches?

64. A pycnometer is a device used to measure density. It weighs 20.455 g empty and 31.486 g when filled with water (d = 1.00 g/cm^3). Pieces of an alloy are put into the empty, dry pycnometer. The mass of the alloy and pycnometer is 28.695 g. Water is added to the alloy to exactly fill the pycnometer. The mass of the pycnometer, water, and alloy is 38.689 g. What is the density of the alloy?

65. Titanium is used in airplane bodies because it is strong and light. It has a density of 4.55 g/cm^3. If a cylinder of titanium is 7.75 cm long and has a mass of 153.2 g, calculate the diameter of the cylinder. ($V = \pi r^2 h$, where V is the volume of the cylinder, r is its radius, and h is the height.)

Conceptual Questions

66. Label each of the properties of iodine as intensive (I) or extensive (E).

(a) Its density is 4.93 g/mL.
(b) It is purple.
(c) It melts at 114°C.
(d) One hundred grams of melted iodine has a volume of 122 mL.
(e) Some iodine crystals are large.

67. How do you distinguish

(a) chemical properties from physical properties?
(b) distillation from filtration?
(c) a solute from a solution?

68. What is the difference between

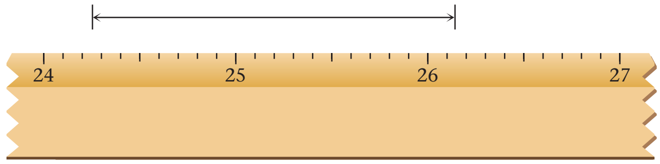
(a) mass and density?
(b) an extensive and an intensive property?
(c) a solvent and a solution?

69. Mercury, ethyl alcohol, and lead are poured into a cylinder. Three distinct layers are formed. The densities of the three substances are

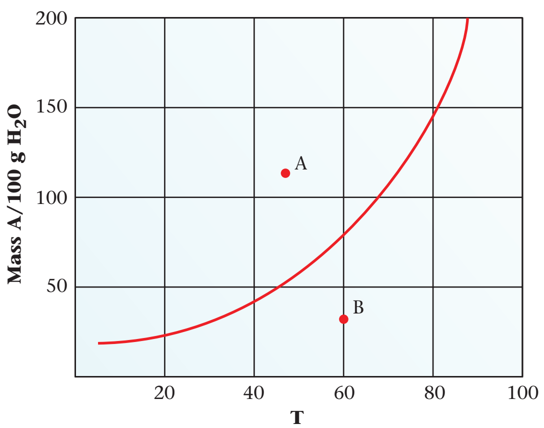
$$\text{mercury} = 13.55 \text{ g/cm}^3$$
$$\text{ethyl alcohol} = 0.78 \text{ g/cm}^3$$
$$\text{lead} = 11.4 \text{ g/cm}^3$$

Sketch the cylinder with the three layers. Identify the substance in each layer.

70. How many significant figures are there in the length of this line?



71. Consider the following solubility graph.

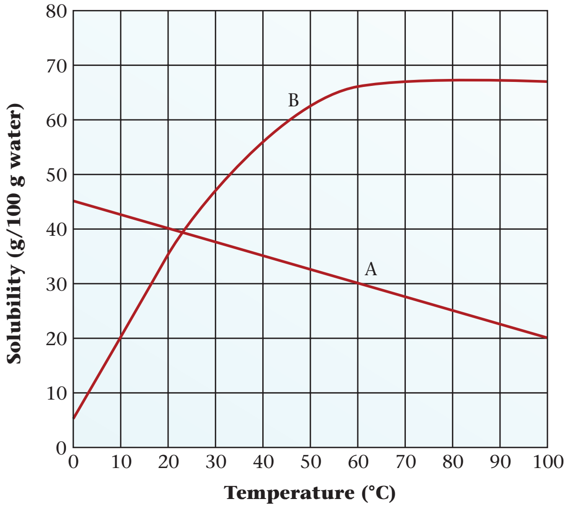


(a) At point A, how many grams of the compound are dissolved in 100 g of water? Is the solution saturated, unsaturated, or supersaturated?
(b) At point B, how many grams of the compound are dissolved in 100 g of water? Is the solution saturated, unsaturated, or supersaturated?
(c) How would you prepare a saturated solution at 30°C?

72. Given the following solubility curves, answer the following questions:

(a) In which of the two compounds can more solute be dissolved in the same amount of water when the temperature is decreased?
(b) At what temperature is the solubility of both compounds the same?

(**c**) Will an increase in temperature always increase solubility of a compound? Explain.



Challenge Problems

73. A different civilization on a distant planet has developed a new temperature scale based on ethyl alcohol. The freezing point of ethyl alcohol (−117°C) is designated as 0°J, and its boiling point (78°C) is designated as 100°J. Assuming that the relationship between °C and °J is linear,

(**a**) draw a graph of the line using the data above.
(**b**) what is the slope of the line?
(**c**) what is the *y*-intercept of the line?
(**d**) Write an equation to convert °J to °C.

74. At what point is the temperature in °F exactly twice that in °C?

75. Oil spreads on water to form a film about 100 nm thick (two significant figures). How many square kilometers of ocean will be covered by the slick formed when one barrel of oil is spilled (1 barrel = 31.5 U.S. gal)?

76. A laboratory experiment requires 12.0 g of aluminum wire ($d = 2.70\ g/cm^3$). The diameter of the wire is 0.200 in. Determine the length of the wire, in centimeters, to be used for this experiment. The volume of a cylinder is $\pi r^2 \ell$, where r = radius and ℓ = length.

77. An average adult breathes about 8.50×10^3 L of air per day. The concentration of lead in highly polluted urban air is 7.0×10^{-6} g of lead per one m^3 of air. Assume that 75% of the lead is present as particles less than 1.0×10^{-6} m in diameter, and that 50% of the particles below that size are retained in the lungs. Calculate the mass of lead absorbed in this manner in 1 year by an average adult living in this environment.

78. A student determines the density of a metal by taking two cylinders of equal mass and volume. Cylinder A is filled with mercury and the metal. The mercury and metal together have a mass of 92.60 g. Cylinder B is filled only with mercury. The mass of the mercury in cylinder B is 52.6 g more than the mass of mercury and metal in cylinder A. What is the density of the metal? The density of mercury is $13.6\ g/cm^3$.