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INTRODUCTION: MATTER AND MEASUREMENT

The giant elliptical galaxy Centaurus A, 10 to 16 million light-years from Earth, as seen by the Hubble Space Telescope, shows the blue glow of young star clusters among clouds of dust and gas. Compression of hydrogen gas clouds causes new star formation, seen as red patches in the image.

KEY CONCEPTS

1.1 THE STUDY OF CHEMISTRY
We begin with a brief description of what chemistry is and why it is useful to learn chemistry.

1.2 CLASSIFICATIONS OF MATTER
We discuss some fundamental ways of classifying matter, distinguishing between *pure substances* and *mixtures* and between *elements* and *compounds*.

1.3 PROPERTIES OF MATTER
We describe the different characteristics or *properties* of matter, used to characterise, identify and separate substances.

1.4 UNITS OF MEASUREMENT
We observe that many properties of matter rely on quantitative measurements involving numbers and units that are based on the *metric system*.

1.5 UNCERTAINTY IN MEASUREMENT
We observe that all measured quantities have an inherent uncertainty that is expressed by the number of *significant figures* used to report the quantity. Significant figures are also used to express the uncertainty associated with calculations involving measured quantities.

The universe is full of mysteries that we will probably never comprehend. And even on Earth some of the simple things that we see and experience can be quite mysterious. How do we obtain electricity from a battery? How does a plant grow? How does a modern LED television screen work? There are innumerable questions, which while seemingly unanswerable, can actually be answered by the study of chemistry.

Chemistry is the study of the properties of matter and the changes that matter undergoes.

This first chapter lays a foundation for our studies by providing an overview of what chemistry is about and what chemists do. The preceding Key Concepts list indicates the chapter organisation and some of the ideas that we will consider.

1.1 THE STUDY OF CHEMISTRY

The Atomic and Molecular Perspective of Chemistry

Chemistry involves studying the properties and behaviour of matter. **Matter** is the *physical* material of the universe: it is anything that has mass and occupies space. A **property** is any characteristic that allows us to recognise a particular type of matter and to distinguish it from other types. This book, your body, the clothes you are wearing, the water you drink and the air you are breathing are all examples of matter. It has long been known that all matter is composed of infinitesimally small building blocks called **atoms**. Despite the tremendous variety of matter in the universe, there are only about 100 different types of atoms that occur in nature and these, combined in various combinations and proportions, constitute all of the matter of the universe. We will see that the properties of matter relate not only to the kinds of atoms it contains (*composition*), but also to the arrangements of these atoms (*structure*).

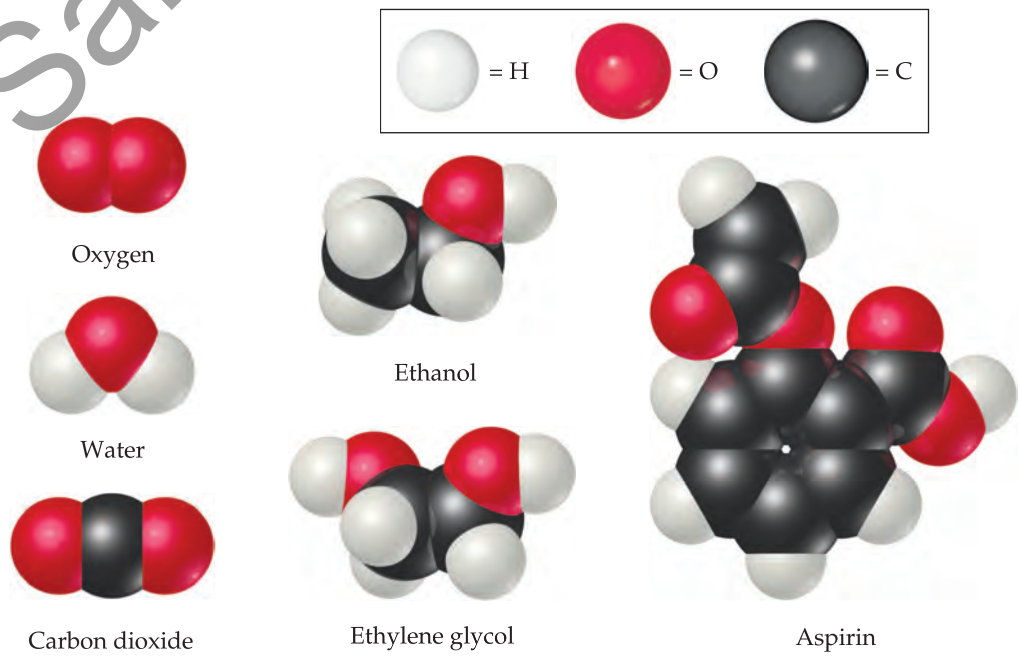
Atoms can combine to form **molecules** in which two or more atoms are joined in specific shapes. Throughout this text you will see molecules represented using coloured spheres to show how their component atoms connect to each other (▼ **FIGURE 1.1**). The colour merely provides a convenient way to distinguish between different kinds of atoms. As examples, compare the molecules of ethanol and ethylene glycol, depicted in Figure 1.1. Notice that these molecules differ somewhat in composition. Ethanol contains one red sphere, which represents one oxygen atom, whereas ethylene glycol contains two.

Even apparently minor differences in the composition or structure of molecules can cause profound differences in their properties. Ethanol, also called grain alcohol, is the alcohol in beverages such as beer and wine. Ethylene glycol, however, is a viscous liquid used as coolant in car radiators.

Every change in the observable world—from boiling water to the changes that occur as our bodies combat invading viruses—has its basis in the world of

FIGURE IT OUT

How many carbon atoms are in one aspirin molecule?



▶ **FIGURE 1.1 Molecular models.** The white, black and red spheres represent atoms of hydrogen, carbon and oxygen, respectively.

atoms and molecules. Thus as we proceed with our study of chemistry, we will find ourselves thinking in two realms: the *macroscopic* realm of ordinary-sized objects (*macro* = large) and the *submicroscopic* realm of atoms and molecules. We make our observations in the macroscopic world—in the laboratory and in our everyday surroundings. In order to understand that world, however, we must visualise how atoms and molecules behave at the submicroscopic level. Chemistry is the science that seeks to understand the properties and behaviour of matter by studying the properties and behaviour of atoms and molecules.

CONCEPT CHECK 1

a. In round numbers, about how many elements are there?
b. What submicroscopic particles are the building blocks of matter?

Why Study Chemistry?

You will note, when studying any scientific discipline, whether it be biology, engineering, medicine, agriculture, geology and so forth, that chemistry is an integral part of your curriculum. This is because chemistry, by its very nature, is the *central science*, central to a fundamental understanding of other sciences and technologies. Chemistry provides an important understanding of our world and how it works. It is an extremely practical science that greatly impacts on our daily living. Indeed, chemistry lies near the heart of many matters of public concern: improvement of health care, conservation of natural resources, protection of the environment and provision of our everyday needs for food, clothing and shelter.

Using chemistry, we have discovered pharmaceutical chemicals that enhance our health and prolong our lives. We have increased food production through the development of fertilisers and pesticides. We have developed plastics and other materials that are used in almost every facet of our lives. Unfortunately, some chemicals also have the potential to harm our health or the environment. It is in our best interests as educated citizens and consumers to understand the profound effects, both positive and negative, that chemicals have on our lives and to strike an informed balance about their uses.

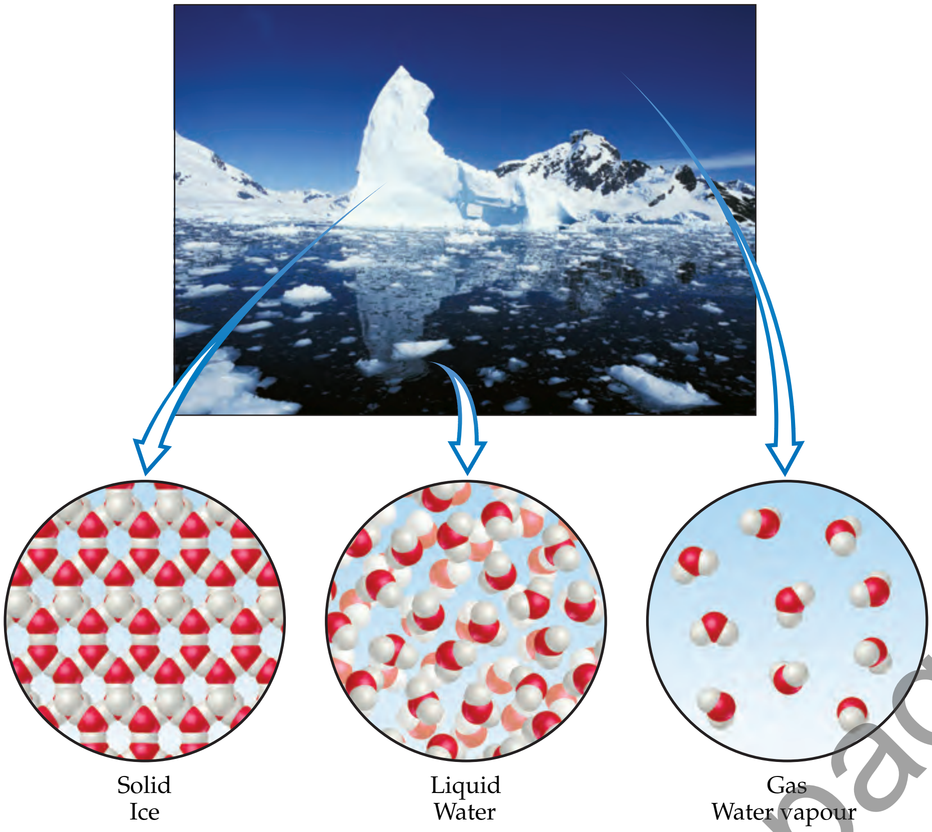
1.2 CLASSIFICATIONS OF MATTER

Let's begin our study of chemistry by examining some fundamental ways in which matter is classified and described. Two principal ways of classifying matter are according to its physical state (gas, liquid or solid) and according to its composition (element, compound or mixture) as explained below.

States of Matter

A sample of matter can be a gas, a liquid or a solid. These three forms of matter are called the **states of matter**. The states of matter differ in some of their simple observable properties. A **gas** (also known as *vapour*) has no fixed volume or shape; rather, it conforms to the volume and shape of its container. A gas can be compressed to occupy a smaller volume, or it can expand to occupy a larger one. A **liquid** has a distinct volume independent of its container but has no specific shape: it assumes the shape of the portion of the container that it occupies. A **solid** has both a definite shape and a definite volume. Neither liquids nor solids can be compressed to any appreciable extent.

The properties of the states can be understood on the molecular level (▼ **FIGURE 1.2**). In a gas the molecules are far apart and are moving at high speeds, colliding repeatedly with each other and with the walls of the container. In a liquid the molecules are packed more closely together, but still move rapidly, allowing them to slide over each other; thus liquids pour easily. In a solid the molecules are held tightly together, usually in definite arrangements, so



▲ **FIGURE 1.2 The three physical states of water: water vapour, liquid water and ice.** Here we see both the liquid and solid states of water. We cannot see water vapour. What we see when we look at steam or clouds is tiny droplets of liquid water dispersed in the atmosphere. The molecular views show that the molecules in the solid are arranged in a more orderly way than in the liquid. The molecules in the gas are much further apart than those in the liquid or the solid.

the molecules can wiggle only slightly in their otherwise fixed positions. Changes in temperature and/or pressure can lead to a conversion from one state of matter to another, illustrated by such familiar processes as ice melting or water evaporating.

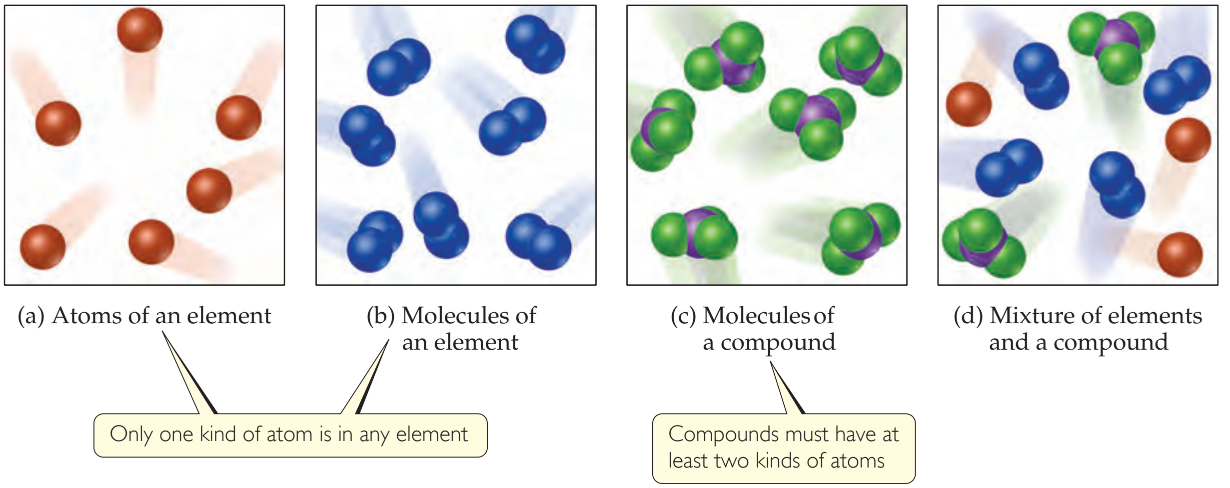
Composition of Matter

When we discuss matter in daily language we often use the word substance as in, 'This is a peculiar substance!' In fact, the word substance is used in everyday language as a substitute for matter which may be one kind of matter or a mixture of more than one kind of matter. In chemistry, however, the word substance means *matter of uniform composition throughout a sample*, as well as having distinct properties. To emphasise this, we usually use the term **pure substance**. However, even when we use the word substance by itself, it is understood to refer to a pure form of matter. For example, oxygen, water, table sugar (sucrose), table salt (sodium chloride) should be referred to as pure substances but more usually are referred to simply as substances.

All substances are either elements or compounds. **Elements** cannot be decomposed into simpler substances; they may be atoms, or molecules composed of only one kind of atom (▼ **FIGURE 1.3(a), (b)**). **Compounds** are substances composed of two or more different elements, so they contain two or more kinds of atoms (Figure 1.3(c)). Water, for example, is a compound composed of two elements, hydrogen and oxygen. Figure 1.3(d) shows a mixture of substances. **Mixtures** are combinations of two or more substances in which each

FIGURE IT OUT

How do the molecules of a compound differ from the molecules of an element?



▲ **FIGURE 1.3 Molecular comparison of elements, compounds and mixtures.**

substance retains its own chemical identity and which can be separated into the individual pure substances by various means.

Some of the more common elements are listed in ▼ TABLE 1.1, along with the chemical abbreviations of their names—**chemical symbols**—used to denote them. All the known elements and their symbols are listed on the inside front cover of this text. The table in which the symbol for each element is enclosed in a box is called the periodic table which is discussed later (Section 2.5, 'The Periodic Table').

Find out more on page 39

The symbol for each element consists of one or two letters, with the first letter capitalised. These symbols are often derived from the English name for the element, but sometimes they are derived from a foreign name (usually Latin) instead (last column in Table 1.1). You will need to know these symbols and to learn others as we encounter them in the text.

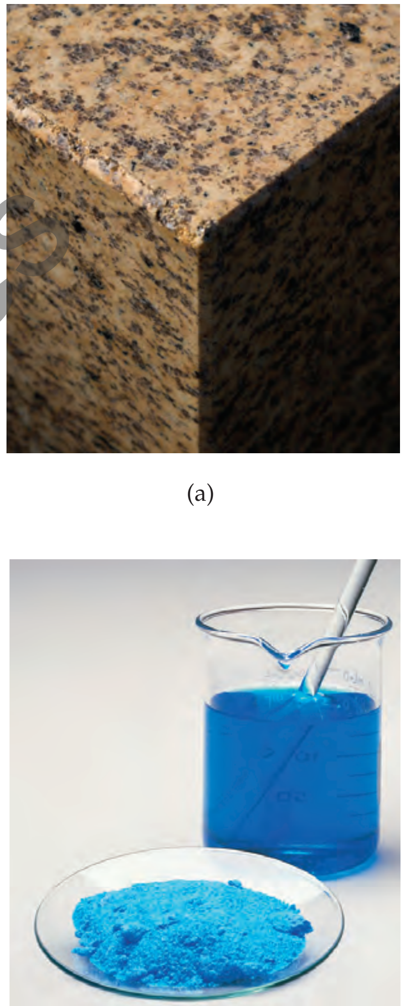
The observation that the elemental composition of a pure compound is always the same is known as the **law of constant composition** (or the **law of definite proportions**). It was first put forth by the French chemist Joseph Louis Proust (1754–1826) in about 1800. Although this law has been known for 200 years, the general belief persists among some people that a fundamental difference exists between compounds prepared in the laboratory and the corresponding compounds found in nature. This is not true: a pure compound has the same composition and properties regardless of its source. Both chemists and nature must use the same elements and operate under the same natural laws to form compounds.

CONCEPT CHECK 2

Hydrogen, oxygen and water are all composed of molecules. What is it about the molecules of water that makes water a compound?

Most of the matter we encounter consists of mixtures of different substances. Each substance in a mixture retains its own chemical identity and hence its own properties. Whereas pure substances have fixed compositions, the compositions of mixtures can vary. A cup of sweetened coffee, for example, can contain either a little sugar or a lot. The substances making up a mixture (such as sugar and water) are called *components* of the mixture.

Some mixtures do not have the same composition, properties and appearance throughout. Both rocks and wood, for example, vary in texture and appearance throughout any typical sample. Such mixtures are *heterogeneous* (▶ FIGURE 1.4(a)). Mixtures that are uniform throughout are *homogeneous*. Air is a homogeneous mixture of the gaseous substances nitrogen, oxygen and smaller amounts of other substances. The nitrogen in air has all the properties that pure nitrogen does because both the pure substance and the mixture contain the same nitrogen molecules. Salt, sugar and many other substances dissolve in water to form homogeneous mixtures (Figure 1.4(b)). Homogeneous mixtures are also called **solutions**.



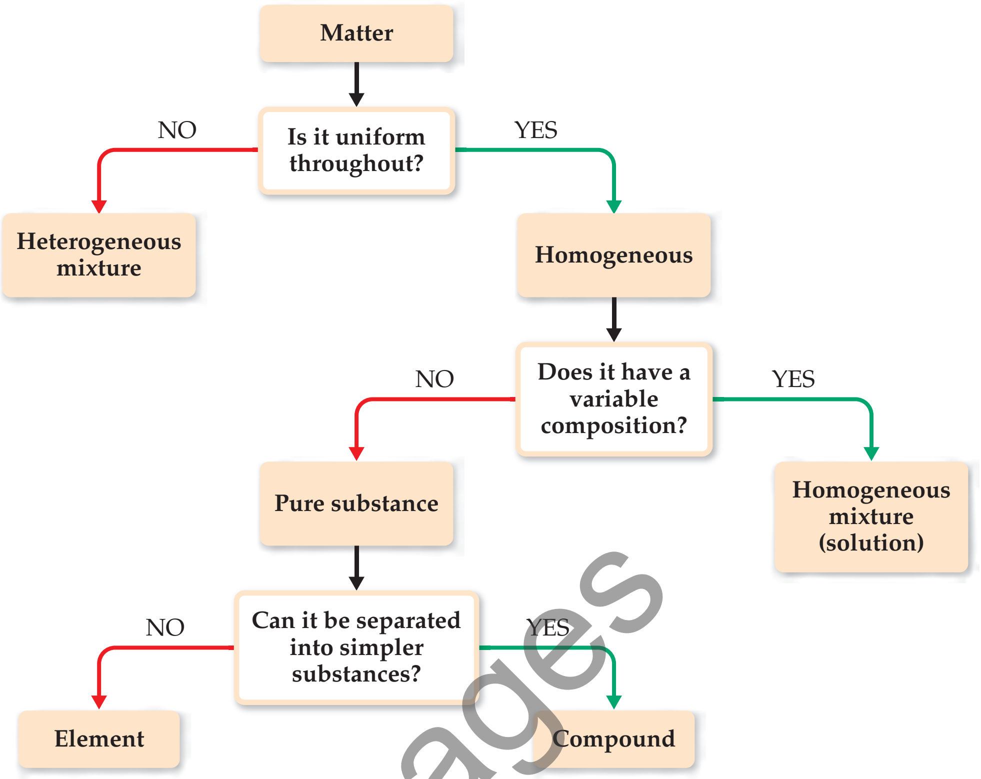
(a)

(b)

▲ **FIGURE 1.4** **Mixtures.** (a) Many common materials, including rocks, are heterogeneous. This granite shows a heterogeneous mixture of silicon dioxide and other metal oxides. (b) Homogeneous mixtures are called solutions. Many substances, including the blue solid shown in this photo (copper sulfate), dissolve in water to form solutions.

TABLE 1.1 • Some common elements and their symbols

Carbon	C	Aluminium	Al	Copper	Cu (from *cuprum*)
Fluorine	F	Bromine	Br	Iron	Fe (from *ferrum*)
Hydrogen	H	Calcium	Ca	Lead	Pb (from *plumbum*)
Iodine	I	Chlorine	Cl	Mercury	Hg (from *hydrargyrum*)
Nitrogen	N	Helium	He	Potassium	K (from *kalium*)
Oxygen	O	Lithium	Li	Silver	Ag (from *argentum*)
Phosphorus	P	Magnesium	Mg	Sodium	Na (from *natrium*)
Sulfur	S	Silicon	Si	Tin	Sn (from *stannum*)



▶ **FIGURE 1.5 Classification of matter.** All matter is classified ultimately as either an element or a compound.

▲ **FIGURE 1.5** summarises the classification of matter into elements, compounds and mixtures.

SAMPLE EXERCISE 1.1 Distinguishing between elements, compounds and mixtures

'White gold', used in jewellery, contains two elements: gold and a 'white' metal such as palladium. Two different samples of white gold differ in the relative amounts of gold and palladium that they contain. Both samples are uniform in composition throughout. Without knowing any more about the materials, use Figure 1.5 to characterise and classify white gold.

SOLUTION

Because the material is uniform throughout, it is homogeneous. Because its composition differs for the two samples, it cannot be a compound. Instead, it must be a homogeneous mixture. Gold and palladium can be said to form a solid solution with one another.

PRACTICE EXERCISE

All aspirin is composed of 60.0% carbon, 4.5% hydrogen and 35.5% oxygen by mass, regardless of its source. Use Figure 1.5 to characterise and classify aspirin.

Answer: It is a compound because it has constant composition and can be separated into several elements.

(See also Exercises 1.6, 1.7, 1.31.)

1.3 PROPERTIES OF MATTER

Every substance has a unique set of properties. The properties of matter can be categorised as physical or chemical. **Physical properties** can be measured without changing the identity and composition of the substance. These properties include colour, odour, density, melting point, boiling point and hardness.

Chemical properties describe the way a substance may change, or *react*, to form other substances. A common chemical property is flammability, the ability of a substance to burn in the presence of oxygen.

Some properties—such as temperature, melting point and density—do not depend on the amount of the sample being examined. These properties, called **intensive properties**, are particularly useful in chemistry because many can be used to *identify* substances. **Extensive properties** of substances depend on the quantity of the sample, with two examples being mass and volume. Extensive properties relate to the *amount* of substance present.

Physical and Chemical Changes

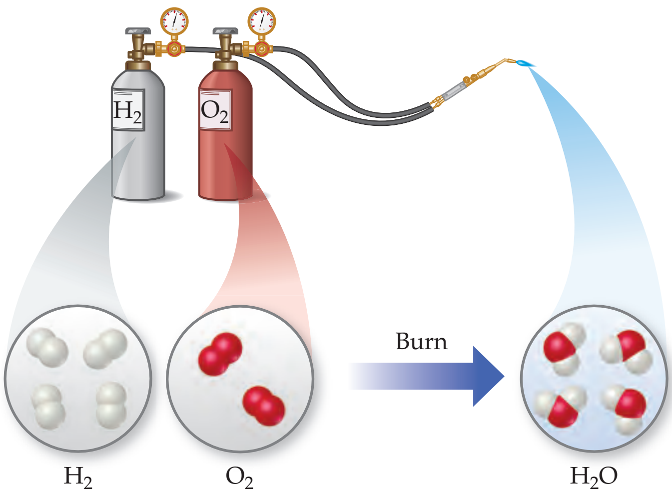
As with the properties of a substance, the changes that substances undergo can be classified as either physical or chemical. During **physical changes** a substance changes its physical appearance, but not its composition. The evaporation of water is a physical change. When water evaporates, it changes from the liquid state to the gas state, but it is still composed of water molecules, as depicted earlier in Figure 1.2. All **changes of state** (for example, from liquid to gas or from liquid to solid) are physical changes.

In **chemical changes** (also called **chemical reactions**) a substance is transformed into a chemically different substance. When hydrogen burns in air, for example, it undergoes a chemical change because it combines with oxygen to form water. The molecular-level view of this process is depicted in ▼ **FIGURE 1.6**.

> **CONCEPT CHECK 3**
>
> Which of the following is a physical change, and which is a chemical change? Explain.
>
> **a.** Plants use carbon dioxide and water to make sugar.
>
> **b.** Water vapour in the air on a cold day forms frost.

Separation of Mixtures

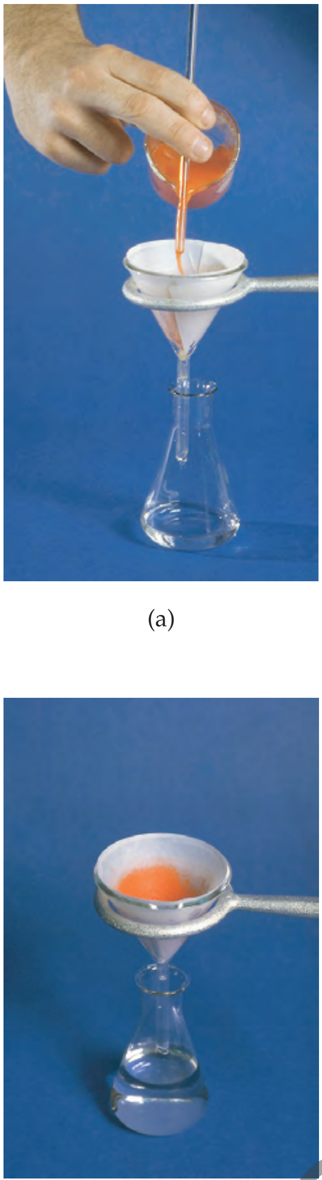
Because each component of a mixture retains its own properties, we can separate a mixture into its components by taking advantage of the differences in their properties. For example, a heterogeneous mixture of iron filings and gold filings could be sorted individually by colour into iron and gold. A less tedious approach would be to use a magnet to attract the iron filings, leaving the gold ones behind. We can also take advantage of an important chemical difference



▲ **FIGURE 1.6** **A chemical reaction.**

between these two metals: many acids dissolve iron but not gold. Thus if we put our mixture into an appropriate acid, the iron would dissolve and the gold would be left behind. The two could then be separated by *filtration*, a procedure illustrated in ◀ **FIGURE 1.7**. We would have to use other chemical reactions, which we will learn about later, to transform the dissolved iron back into metal.

An important method of separating the components of a homogeneous mixture is *distillation*, a process that depends on the different abilities of substances to form gases. For example, if we boil a solution of salt and water, the water evaporates, forming a gas, and the salt is left behind. The gaseous water can be converted back to a liquid on the walls of a condenser, as shown in the apparatus depicted in ▼ **FIGURE 1.8**.



(a)

(b)

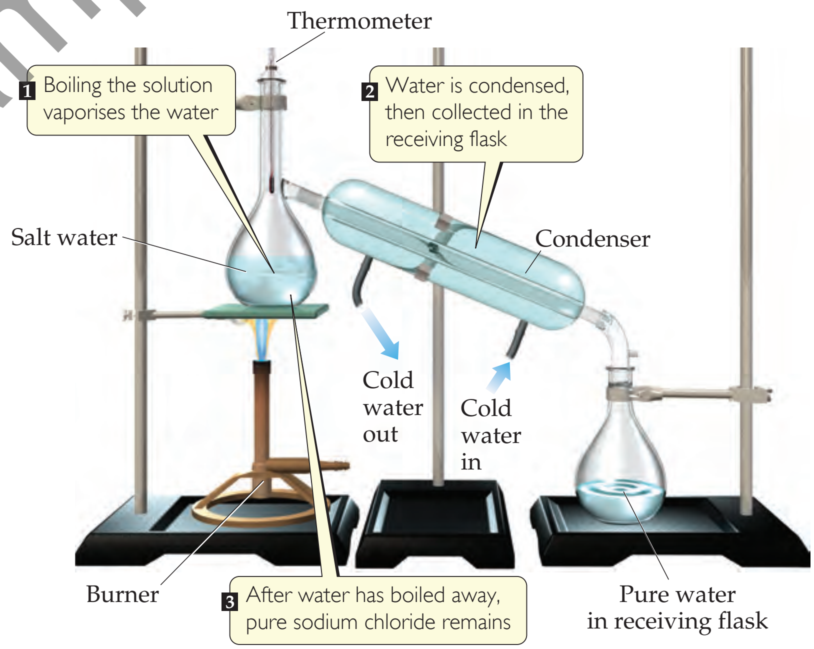
▲ **FIGURE 1.7 Separation by filtration.** (a) A mixture of a solid and a liquid is poured through a porous medium, in this case filter paper. (b) The liquid has passed through the paper but the solid remains on the paper.

1.4 UNITS OF MEASUREMENT

Science depends on making accurate measurements of the phenomena being observed. A measurement consists of a number and a scale, which is referred to as a unit. When a number represents a measured quantity, the units of that quantity must be specified. To say that the length of a pencil is 17.5 is meaningless. Expressing the number with its units, 17.5 centimetres (cm), properly specifies the length. The units used for scientific measurements are those of the **metric system**, which was first developed in France in the eighteenth century. Most countries in the world also use this system in their daily life.

SI Units

In 1960 an international agreement was reached specifying a particular choice of metric units for use in scientific measurements. These preferred units are called **SI units**, after the French *Système International d'Unités*. This system has seven *base units* from which all other units are derived. ▶ **TABLE 1.2** lists these base units and their symbols. In this chapter we consider the base units for length, mass and temperature.



▲ **FIGURE 1.8 Distillation.** A simple apparatus for the separation of a sodium chloride solution (salt water) into its components. Boiling the solution evaporates the water, which is condensed then collected in the receiving flask. After all the water has boiled away, pure sodium chloride remains in the boiling flask.

A CLOSER LOOK

THE SCIENTIFIC METHOD

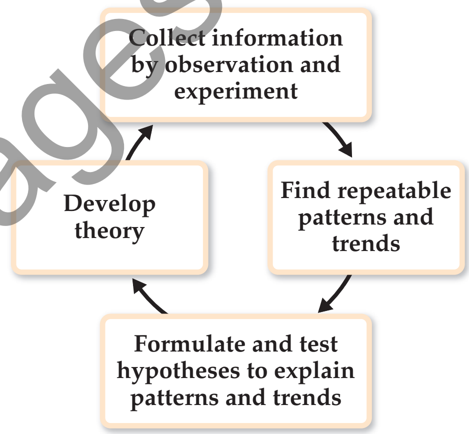
Although two scientists rarely approach the same problem in exactly the same way, there are guidelines for the practice of science that have come to be known as the scientific method. These guidelines are outlined in ▶ FIGURE 1.9. We begin by collecting information, or *data*, by observation and experiment. The collection of information, however, is not the ultimate goal. The goal is to find a pattern or sense of order in our observations and to understand the origin of this order.

As we perform our experiments, we may begin to see patterns that lead us to a *tentative explanation*, or hypothesis, that guides us in planning further experiments. A key feature of a good hypothesis is that it proposes a mechanism and can be used to make predictions about new experiments. If a hypothesis is sufficiently general and is continually effective in predicting facts yet to be observed, it is called a theory. A theory is *an explanation of the general causes of certain phenomena, with considerable evidence or facts to support it*. For example, Einstein's theory of relativity was a revolutionary new way of thinking about space and time. It was more than just a simple hypothesis, however, because it could be used to make predictions that could be tested experimentally. The results of these experiments were generally in agreement with Einstein's predictions and were not explainable by earlier theories. Despite the landmark achievements of Einstein's theory, scientists can never say the theory is proven. A theory that has excellent predictive power today may not work as well in the future as more data are collected and improved scientific equipment is developed. Thus science is always a work in progress.

Eventually, we may be able to tie together a great number of observations in a single statement or equation called a scientific law. A scientific law is *a concise verbal statement or a mathematical equation that summarises a broad variety of observations and experiences*. We tend to think of the laws of nature as the basic rules under which nature operates. However, it is not so much that matter obeys the laws of nature, but rather that the laws of nature describe the behaviour of matter.

As we proceed through this text, we will rarely have the opportunity to discuss the doubts, conflicts, clashes of personalities and revolutions of perception that have led to our present ideas. We need to be aware that just because we can spell out the results of science so concisely and neatly in textbooks does not mean that scientific progress is smooth, certain and predictable. Some of the ideas we present in this text took centuries to develop and involved large numbers of scientists. We gain our view of the natural world by standing on the shoulders of the scientists who came before us. Take advantage of this view. As you study, exercise your imagination. Don't be afraid to ask daring questions when they occur to you. You may be fascinated by what you discover!

RELATED EXERCISE: 1.32



▲ FIGURE 1.9 **The scientific method.** The scientific method is a general approach to problems that involves making observations, confirming that they are reproducible, seeking patterns in the observations, formulating hypotheses to explain the observations and testing these hypotheses by further experiments. Those hypotheses that withstand such tests and prove themselves useful in explaining and predicting behaviour become known as theories.

TABLE 1.2 • SI base units

Physical quantity	Name of unit	Abbreviation
Mass	Kilogram	kg
Length	Metre	m
Time	Second	s
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

TABLE 1.3 • Selected prefixes used in the metric system

Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 gigametre (Gm) = 1×10^9 m
Mega	M	10^6	1 megametre (Mm) = 1×10^6 m
Kilo	k	10^3	1 kilometre (km) = 1×10^3 m
Deci	d	10^{-1}	1 decimetre (dm) = 1×10^{-1} m
Centi	c	10^{-2}	1 centimetre (cm) = 1×10^{-2} m
Milli	m	10^{-3}	1 millimetre (mm) = 1×10^{-3} m
Micro	μ[a]	10^{-6}	1 micrometre (μm) = 1×10^{-6} m
Nano	n	10^{-9}	1 nanometre (nm) = 1×10^{-9} m
Pico	p	10^{-12}	1 picometre (pm) = 1×10^{-12} m
Femto	f	10^{-15}	1 femtometre (fm) = 1×10^{-15} m
Atto	a	10^{-18}	1 attometre (am) = 1×10^{-18} m
Zepto	z	10^{-21}	1 zeptometre (zm) = 1×10^{-21} m

[a] This is the Greek letter mu (pronounced 'mew').

Sometimes the base units are not convenient and so prefixes are used to indicate decimal fractions or multiples of various units. For example, the prefix *milli-* represents a 10^{-3} fraction of a unit: a milligram (mg) is 10^{-3} gram (g), a millimetre (mm) is 10^{-3} metre (m), and so forth. ▲ TABLE 1.3 presents the prefixes commonly encountered in chemistry. In using SI units and in working problems throughout this text, you must be comfortable using exponential notation. If you are unfamiliar with exponential notation or want to review it, refer to Appendix A.

Although non-SI units are being phased out, there are still some that are commonly used by scientists. Whenever we first encounter a non-SI unit in the text, the proper SI unit will also be given.

CONCEPT CHECK 4

Which of the following quantities is the smallest: 1 mg, 1 μg or 1 pg?

Length and Mass

The SI base unit of *length* is the **metre** (m).

Mass* (*m*) is a measure of the amount of material in an object. The SI base unit of mass is the **kilogram** (kg). This base unit is unusual because it uses a prefix, *kilo-*, instead of the word *gram* alone. We obtain other units for mass by adding prefixes to the word *gram*.

SAMPLE EXERCISE 1.2 Using metric prefixes

What is the name given to the unit that equals **(a)** 10^{-9} gram, **(b)** 10^{-6} second, **(c)** 10^{-3} metre?

SOLUTION

In each case we can refer to Table 1.3, finding the prefix related to each of the decimal fractions: **(a)** nanogram, ng, **(b)** microsecond, μs, **(c)** millimetre, mm.

* Mass and weight are not interchangeable terms and are often incorrectly thought to be the same. The weight of an object is the force that its mass exerts due to gravity. In space, where gravitational forces are very weak, astronauts can be weightless, but they cannot be massless. In fact, an astronaut's mass in space is the same as it is on Earth.

PRACTICE EXERCISE

(a) What decimal fraction of a second is a picosecond, ps? **(b)** Express the measurement 6.0×10^3 m using a prefix to replace the power of ten. **(c)** Use exponential notation to express 3.76 mg in grams.

Answers: **(a)** 10^{-12} second, **(b)** 6.0 km, **(c)** 3.76×10^{-3} g

(See also Exercises 1.17, 1.18.)

Temperature

Temperature is a measure of the hotness or coldness of an object. Indeed, temperature is a physical property that determines the direction of heat flow. Heat always flows spontaneously from a substance at higher temperature to one at lower temperature. Thus we feel the influx of heat when we touch a hot object, and we know that the object is at a higher temperature than our hand.

The temperature scales commonly employed in scientific studies are the Celsius and Kelvin scales. The **Celsius scale** is also the everyday scale of temperature in most countries (▼ **FIGURE 1.10**). It was originally based on the assignment of 0 °C to the freezing point of water and 100 °C to its boiling point at sea level (Figure 1.10).

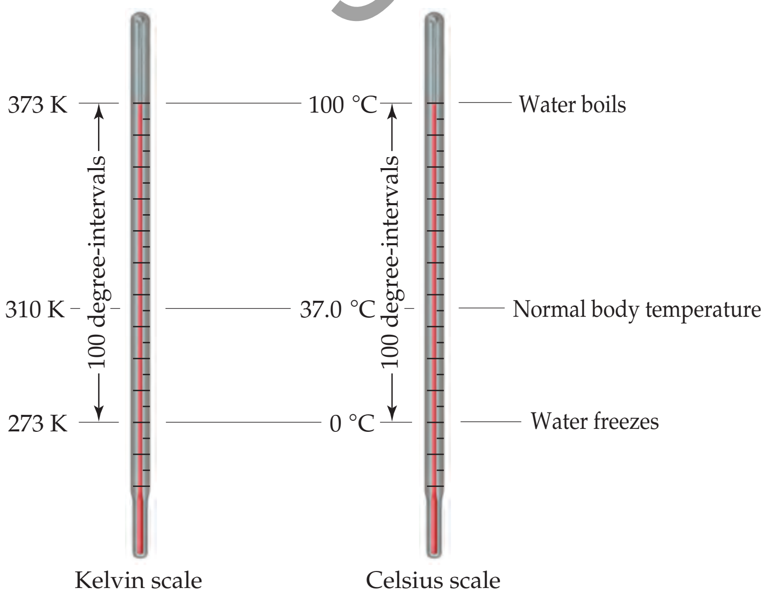
The **Kelvin scale** is the SI temperature scale, and the SI unit of temperature is the kelvin (K). Historically, the Kelvin scale was based on the properties of gases; its origins are considered in Chapter 10. Zero on this scale is the lowest attainable temperature, –273.15 °C, a temperature referred to as *absolute zero*. Both the Celsius and Kelvin scales have equal-sized units—that is, a kelvin is the same size as a degree Celsius. Thus the Kelvin and Celsius scales are related as follows:

$$K = °C + 273.15 \qquad [1.1]$$

The freezing point of water, 0 °C, is 273.15 K (Figure 1.10). Notice that we do not use a degree sign (°) with temperatures on the Kelvin scale.

FIGURE IT OUT

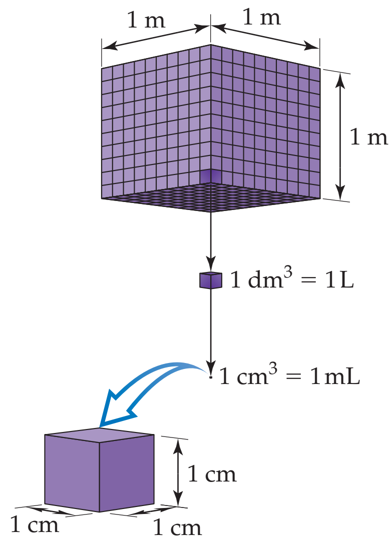
True or false: The 'size' of a degree on the Celsius scale is the same as the 'size' of a degree on the Kelvin scale.



◀ **FIGURE 1.10 Comparison of the Kelvin and Celsius temperature scales.**

FIGURE IT OUT

How many 1 dm^3 bottles are required to contain 1 m^3 of liquid?



▲ **FIGURE 1.11 Volume relationships.** The volume occupied by a cube 1 m on each edge is one cubic metre, 1 m^3. Each cubic metre contains 1000 dm^3. One litre is the same volume as one cubic decimetre, 1 L = 1 dm^3. Each cubic decimetre contains 1000 cubic centimetres, 1 dm^3 = 1000 cm^3. One cubic centimetre equals one millilitre, 1 cm^3 = 1 mL.

SAMPLE EXERCISE 1.3 Converting units of temperature

If a weather forecaster predicts that the temperature for the day will reach 31 °C, what is the predicted temperature in K?

SOLUTION

Referring to Equation 1.1, we have K = 31 + 273 = 304 K.

PRACTICE EXERCISE

Ethylene glycol, the major ingredient in antifreeze, freezes at –11.5 °C. What is the freezing point in K?

Answer: 261.7 K

Derived SI Units

The SI base units in Table 1.2 are used to derive the units of other quantities. To do so, we use the defining equation for the quantity, substituting the appropriate base units. For example, speed is defined as the ratio of distance travelled to elapsed time. Thus the SI unit for speed is the SI unit for distance (length) divided by the SI unit for time, m s^{-1}, which we read as 'metres per second'. We will encounter many derived units, such as those for force, pressure and energy, later in this text. In this chapter we examine the derived units for volume and density.

Volume

The **volume** (V) of a cube is given by its length cubed $(\text{length})^3$. Thus the SI unit of volume is the SI unit of length raised to the third power. The cubic metre, or m^3, is the volume of a cube that is 1 m on each edge. Smaller units are usually used in chemistry. These are the cubic decimetre (dm^3), often referred to as the litre (L), and the cubic centimetre (cm^3), frequently referred to as the millilitre (mL). We will frequently use the terms litre and millilitre in this text because of common usage, but when the abbreviations are written, they will appear as dm^3 and cm^3 (◀ FIGURE 1.11). The relationship between the cubic decimetre and the cubic centimetre is given by:

$$1\ dm^3 = (1\ dm) \times (1\ dm) \times (1\ dm) = (10\ cm) \times (10\ cm) \times (10\ cm) = 1000\ cm^3$$

The devices used most frequently in chemistry to measure volume are illustrated in ▶ FIGURE 1.12. Syringes, burettes and pipettes deliver liquids with more precision than graduated cylinders. Volumetric flasks are used to contain specific volumes of liquid.

TABLE 1.4 • Densities of some selected substances at 25 °C

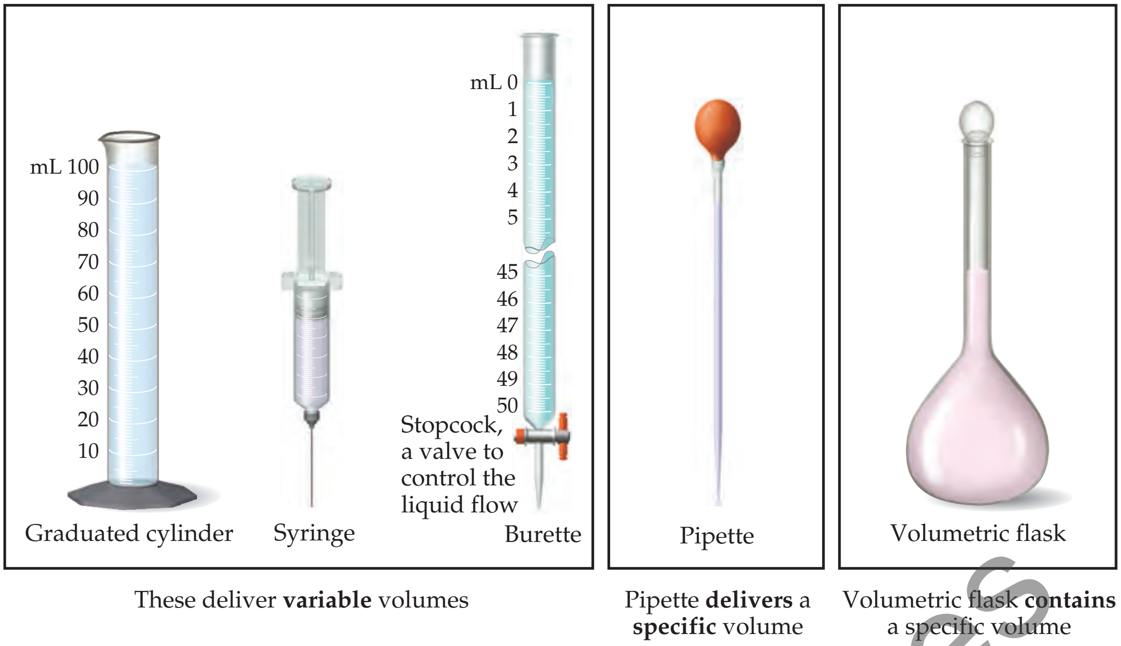
Substance	Density (g cm^{-3})
Air	0.001
Balsa wood	0.16
Ethanol	0.79
Water	1.00
Ethylene glycol	1.09
Table sugar	1.59
Table salt	2.16
Iron	7.9
Gold	19.32

Density

Density (ρ) is a property of matter that is widely used to characterise substances. It is defined as the amount of mass in a unit volume of the substance:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \qquad [1.2]$$

The densities of solids and liquids are commonly expressed in units of grams per cubic centimetre (g cm^{-3}). The densities of some common substances are listed in ◀ TABLE 1.4. It is no coincidence that the density of water is 1.00 g cm^{-3}; the gram was originally defined as the mass of 1 cm^3 of water at a specific temperature. Because most substances change volume when heated or cooled, densities are temperature dependent. When reporting densities, the temperature should be specified. If no temperature is reported, we usually assume that the temperature is 25 °C, close to normal room temperature.



▲ **FIGURE 1.12** **Common volumetric glassware.**

SAMPLE EXERCISE 1.4 Determining density and using density to determine volume or mass

(a) Calculate the density of mercury if 1.00×10^2 g occupies a volume of 7.36 cm^3.
(b) Calculate the volume of 65.0 g of the liquid methanol (wood alcohol) if its density is 0.791 g cm^{-3}.
(c) What is the mass in grams of a cube of gold (density = 19.32 g cm^{-3}) if the length of the cube is 2.00 cm?

SOLUTION

(a) We are given mass and volume, so Equation 1.2 yields

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{1.00 \times 10^2 \text{ g}}{7.36 \text{ cm}^3} = 13.6 \text{ g cm}^{-3}$$

(b) Solving Equation 1.2 for volume and then using the given mass and density gives

$$\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{65.0 \text{ g}}{0.791 \text{ g cm}^{-3}} = 82.2 \text{ cm}^3$$

(c) We can calculate the mass from the volume of the cube and its density. The volume of a cube is given by its length cubed

$$\text{Volume} = (2.00 \text{ cm})^3 = (2.00)^3 \text{ cm}^3 = 8.00 \text{ cm}^3$$

Solving Equation 1.2 for mass and substituting the volume and density of the cube, we have

$$\text{Mass} = \text{volume} \times \text{density} = (8.00 \text{ cm}^3)(19.32 \text{ g cm}^{-3}) = 155 \text{ g}$$

PRACTICE EXERCISE

(a) Calculate the density of a 374.5 g sample of copper if it has a volume of 41.8 cm^3.
(b) A student needs 15.0 g of ethanol for an experiment. If the density of ethanol is 0.789 g cm^{-3}, how many millilitres of ethanol are needed?
(c) What is the mass, in grams, of 25.0 cm^3 of mercury?

Answers: **(a)** 8.96 g/cm^3, **(b)** 19.0 cm^3, **(c)** 340 g

(See also Exercises 1.19–1.21, 1.34, 1.35.)

1.5 UNCERTAINTY IN MEASUREMENT

There are two kinds of numbers in scientific work: *exact numbers* (those whose values are known exactly) and *inexact numbers* (those whose values have some uncertainty). Most of the exact numbers that we will encounter in this course have defined values. For example, there are exactly 12 eggs in a dozen and exactly 1000 g in a kilogram. The number 1 in any conversion factor between units, as in 1 m = 100 cm or 1 kg = 10^6 mg, is also an exact number. Exact numbers can also result from counting numbers of objects. For example, we can count the exact number of marbles in a jar or the exact number of people in a classroom.

Numbers obtained by measurement are always *inexact*. There are always inherent limitations in the equipment used to measure quantities (equipment errors), and there are differences in how different people make the same measurement (human errors). Suppose that 10 students with 10 balances are given the same coin and told to determine its mass. The 10 measurements will probably vary slightly from one another for various reasons. The balances might be calibrated slightly differently, and there might be differences in how each student reads the mass from the balance. Remember: *uncertainties always exist in measured quantities*. Counting very large numbers of objects usually has some associated error as well. Consider, for example, how difficult it is to obtain accurate census information for a city or vote counts for an election.

CONCEPT CHECK 5

Which of the following is an inexact quantity?

a. The number of people in your chemistry class.
b. The mass of a coin.
c. The number of grams in a kilogram.

MY WORLD OF CHEMISTRY

CHEMISTRY IN THE NEWS

Chemistry is a very lively, active field of science. Because it is so central to our lives, there are reports on matters of chemical significance in the news nearly every day. Some tell of recent breakthroughs in the development of new pharmaceuticals, materials and processes. Others deal with environmental and public safety issues. As you study chemistry, we hope you will develop the skills to understand the impact of chemistry on your life more effectively. You need these skills to take part in public discussions and debates about matters related to chemistry that affect your community, the nation and the world. By way of examples, here are summaries of a few recent stories in which chemistry plays a role.

Biofuels are coming

The term biofuel describes fuels that are derived from biomass; the two most common biofuels are bioethanol and biodiesel. Currently only a very small proportion of the world's transport fuels is derived from biomass, being estimated at approximately 3%. However, since biofuels are produced from crops such as corn, soybean and sugar cane, it is rapidly being realised that using food crops to produce transport fuel can actually cause food shortages as farmers sell their crops for a better price to the fuel industry. In Australia, the government has legislated that ethanol produced mainly from sugar cane but also from many other sugar-containing plants can be added to petrol with a maximum of 10%. This is the common E10 petrol now sold at most service stations (▼ **FIGURE 1.13**). Biodiesel is made from vegetable/animal fats and recycled oils

◀ **FIGURE 1.13**
E10. A petrol pump dispensing E10 unleaded petrol.

and greases. It is usually used as a diesel additive but is increasingly being used by itself (B100).

Research into biofuel production is being conducted on an increasing scale often using complicated biomatter such as algae and agricultural wastes in an attempt to minimise the use of food crops for biofuel production.

Super batteries

Electric cars are becoming more common but are still held back by the lack of suitable energy sources. A promising recent development is the lithium iron phosphate ($LiFePO_4$, LFP) battery, invented in 1996, which is used in the General Motors 'Volt' hybrid car. It offers a high current rate, a long cycle life of up to 2000 charge cycles and good energy density of 95–140 Wh kg^{-1}. Recharge time is only 2.5–3 hours but work is in progress to reduce this to 1 hour. The strong advantage of this battery over its lithium cobalt oxide or lithium magnesium oxide predecessors is that the materials are cheap and non-toxic. The batteries are very stable and virtually incombustible during charging and discharging since the P–O bonds are very strong so that oxygen is not readily released as can happen with the lithium cobalt oxide batteries. The *Killacycle* electric motorcycle, specifically built for drag racing, uses a 374 volt pack of $LiFePO_4$ batteries weighing 79.4 kg. It achieved a top speed of 274 km h^{-1} in November 2007.

New solar cell design

A new flexible and lightweight solar cell that achieves a high conversion rate of solar energy to electrical energy has been developed in the United States. The cells use micrometre-sized rods of silicon instead of the conventional silicon wafers. Light entering the cells bounces back and forth many times between the rods guided by aluminium nano-particle reflectors until it is absorbed. The claim is that 85% of usable sunlight is absorbed compared with 17% for current commercial panels. In addition the cells are much better at absorbing light in the near-infrared spectrum. Because of their flexibility, they are even suitable for inserting solar-powered devices in clothes. In addition to these inorganic photovoltaic cells, organic photovoltaic cells, which use highly conjugated organic molecules, are being rapidly developed. They are also known as polymer cells depending on the size of the organic molecules that are used to convert solar energy to electrical energy. Although their possible uses are very interesting they do have very low efficiencies, that is, less than 3% compared with the inorganic devices.

New lighting

The city of Sydney has embarked on a project involving the installation of LED (light emitting diode) street lighting across the Sydney CBD (▶ **FIGURE 1.14**). Recently developed LED light bulbs emit as much light as incandescent or fluorescent light bulbs but use a small fraction of the electrical energy. Instead of emitting light from a vacuum as in an incandescent light bulb or a gas as in a compact fluorescent light bulb (CFL), an LED emits light from a solid, that is, a semiconductor, which is made of a positively and a negatively charged component. When an electric charge is applied to the semiconductor it activates the flow of electrons from the negative to the positive layer. These excited electrons emit light of a certain wavelength. The diode material can be varied but is commonly aluminium gallium arsenide (AlGaAs). A new material gallium nitride promises to deliver light bulbs that will be operational for

▲ **FIGURE 1.14** **LED lighting in Sydney.**

100 000 hours. Although the initial costs of the bulbs are much higher than those of conventional lighting, the savings involved in replacement costs and power usage seem to be very significant.

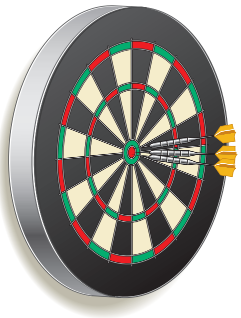
Important antibiotic modified to combat bacterial resistance

Vancomycin is an antibiotic of last resort. It is used only when other antibacterial agents are ineffective. Some bacteria have now developed a resistance to vancomyscin, causing researchers to modify the molecular structure of the substance to make it more effective in killing bacteria. This approach was based on the knowledge that vancomycin works by binding to a protein that is essential to forming bacterial cell walls. Researchers have synthesised a vancomycin analogue in which a CO group has been converted to a CH_2 group (▼ **FIGURE 1.15**). This modification increases the compound's binding affinity in the cell walls of vancomycin-resistant bacteria, making the analogue 100 times more active than vancomycin itself.

▲ **FIGURE 1.15** **Comparing CO and CH_2 groups.** The molecule on the left contains a CO group and the one on the right contains a CH_2 group. This subtle difference is similar to how the much more complex vancomycin molecule was modified.

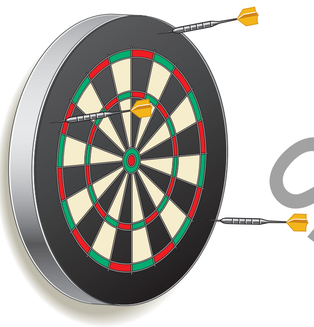
FIGURE IT OUT

How would the darts be positioned on the target for the case of 'good accuracy, poor precision'?



Good accuracy
Good precision

Poor accuracy
Good precision



Poor accuracy
Poor precision

▲ **FIGURE 1.16 Precision and accuracy.** The distribution of darts in a target illustrates the difference between accuracy and precision.

Precision and Accuracy

The terms 'precision' and 'accuracy' are often used in discussing the uncertainties of measured values. **Precision** is a measure of how closely individual measurements agree with one another. **Accuracy** refers to how closely individual measurements agree with the correct, or 'true', value. The analogy of darts stuck in a dartboard pictured in ◀ **FIGURE 1.16** illustrates the difference between these two concepts.

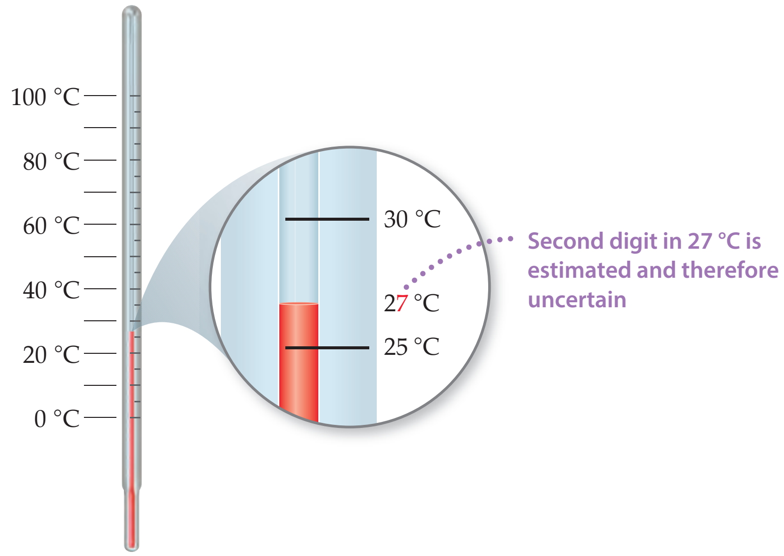
In the laboratory we often perform several different 'trials' of the same experiment. We gain confidence in the accuracy of our measurements if we obtain nearly the same value each time. Figure 1.16 should remind us, however, that precise measurements could be inaccurate. For example, if a very sensitive balance is poorly calibrated, the masses we measure will be consistently either high or low. They will be inaccurate even if they are precise.

Significant Figures

Suppose you determine the mass of a 5 cent piece on a balance capable of measuring to the nearest 0.0001 g. You could report the mass as 2.2405 ± 0.0001 g. The ± notation (read as 'plus or minus') expresses the magnitude of the uncertainty of your measurement. In much scientific work we drop the ± notation with the understanding that there is always some uncertainty in the last digit of the measured quantity. That is, *measured quantities are generally reported in such a way that only the last digit is uncertain*.

▼ **FIGURE 1.17** shows a thermometer with its liquid column between the scale marks. We can read the certain digits from the scale and estimate the uncertain one. From the scale marks on the thermometer, we see that the liquid is between the 25 °C and 30 °C marks. We might estimate the temperature to be 27 °C, being somewhat uncertain of the second digit of our measurement.

All digits of a measured quantity, including the uncertain one, are called **significant figures**. A measured mass reported as 2.2 g has two significant figures, whereas one reported as 2.2405 g has five significant figures. The greater the number of significant figures, the greater is the certainty implied for the measurement. When multiple measurements are made of a quantity, the results can be averaged, and the number of significant figures estimated by using statistical methods.



▲ **FIGURE 1.17 Significant figures in measurements.**

SAMPLE EXERCISE 1.5 **Relating significant figures to the uncertainty of a measurement**

What difference exists between the measured values 4.0 g and 4.00 g?

SOLUTION

Many people would say there is no difference, but a scientist would note the difference in the number of significant figures in the two measurements. The value 4.0 has two significant figures, but 4.00 has three. This difference implies that the first measurement has more uncertainty. A mass of 4.0 g indicates that the uncertainty is in the first decimal place of the measurement. Thus the mass might be anything between 3.9 and 4.1 g, which we can represent as 4.0 ± 0.1 g. A measurement of 4.00 g implies that the uncertainty is in the second decimal place. Thus the mass might be anything between 3.99 and 4.01 g, which we can represent as 4.00 ± 0.01 g. Without further information, we cannot be sure whether the difference in uncertainties of the two measurements reflects the precision or accuracy of the measurement.

PRACTICE EXERCISE

A balance has a precision of $\pm$ 0.001 g. A sample that has a mass of about 25 g is placed on this balance. How many significant figures should be reported for this measurement?

Answer: Five, as in the measurement 24.995 g

To determine the number of significant figures in a reported measurement, read the number from left to right, counting the digits starting with the first digit that is not zero. *In any measurement that is properly reported, all non-zero digits are significant.* Zeros, however, can be used either as part of the measured value or merely to locate the decimal point. Thus zeros may or may not be significant, depending on how they appear in the number. The following guidelines describe the different situations involving zeros.

1. Zeros *between* non-zero digits are always significant—1005 kg (four significant figures); 1.03 cm (three significant figures).
2. Zeros *at the beginning* of a number are never significant; they merely indicate the position of the decimal point—0.02 g (one significant figure); 0.0026 cm (two significant figures).
3. Zeros *at the end* of a number are significant if the number contains a decimal point—0.0200 g (three significant figures); 3.0 cm (two significant figures).

A problem arises when a number ends with zeros but contains no decimal point. In such cases, it is normally assumed that the zeros are not significant. Exponential notation (Appendix A) can be used to indicate clearly whether zeros at the end of a number are significant. For example, a mass of 10 300 g can be written in exponential notation showing three, four or five significant figures depending on how the measurement is obtained:

1.03×10^4 g (three significant figures)
1.030×10^4 g (four significant figures)
1.0300×10^4 g (five significant figures)

In these numbers all the zeros to the right of the decimal point are significant (rules 1 and 3). (The exponential term does not add to the number of significant figures.)

SAMPLE EXERCISE 1.6 **Determining the number of significant figures in a measurement**

How many significant figures are in each of the following numbers (assume that each number is a measured quantity): **(a)** 4.003, **(b)** 6.023×10^{23}, **(c)** 5000?

SOLUTION

(a) Four; the zeros are significant figures. **(b)** Four; the exponential term does not add to the number of significant figures. **(c)** One; we assume that the zeros are not significant when there is no decimal point shown. If the number has more significant figures, it should be written in exponential notation. Thus 5000×10^4 has four significant figures, whereas 5.00×10^3 has three.

PRACTICE EXERCISE

How many significant figures are in each of the following measurements: (a) 3.549 g, **(b)** 2.3×10^4 cm, **(c)** 0.00134 m^3?

Answers: **(a)** Four, **(b)** two, **(c)** three

(See also Exercises 1.24, 1.25.)

Significant Figures in Calculations

When carrying measured quantities through calculations, *the least certain measurement limits the certainty of the calculated quantity and thereby determines the number of significant figures in the final answer*. The final answer should be reported with only one uncertain digit. To keep track of significant figures in calculations, we will make frequent use of two rules, one for multiplication and division and another for addition and subtraction.

1. *For multiplication and division*, the result contains the same number of significant figures as the measurement with the fewest significant figures. When the result contains more than the correct number of significant figures, it must be rounded off. For example, the area of a rectangle whose measured edge lengths are 6.221 cm and 5.2 cm should be reported as 32 cm^2 even though a calculator shows the product of 6.221 and 5.2 to have more digits:

$$\text{Area} = (6.221 \text{ cm})(5.2 \text{ cm}) = 32.3492 \text{ cm}^2 \Rightarrow \text{round off to } 32 \text{ cm}^2$$

   We round off to two significant figures because the least precise number—5.2 cm—has only two significant figures.

2. *For addition and subtraction*, the result has the same number of decimal places as the measurement with the fewest decimal places. Consider the following example in which the uncertain digits appear in colour.

This number limits	20.42	← two decimal places
the number of significant	1.322	← three decimal places
figures in the result →	83.1	← one decimal place
	104.842	← round off to 104.8 (one decimal place)

   We report the result as 104.8 because 83.1 has only one decimal place.

Notice that for multiplication and division, significant figures are counted. For addition and subtraction, decimal places are counted. In determining the final answer for a calculated quantity, exact numbers can be treated as if they have an infinite number of significant figures. This rule applies to many definitions between units. Thus when we say, 'There are 100 centimetres in 1 metre', the number 100 is exact, and we need not worry about the number of significant figures in it.

- In *rounding off* numbers, numbers ending in 0–4 are rounded down and those ending in 5–9 are rounded up. For example, rounding 2.780 to three significant figures would give 2.78 whereas 2.785 would result in 2.79. Similarly, rounding off 2.780 to two significant figures gives 2.8 whereas 4.645 would give 4.6.

SAMPLE EXERCISE 1.7 **Determining the number of significant figures in a calculated quantity**

The width, length and height of a small box are 15.5 cm, 27.3 cm and 5.4 cm, respectively. Calculate the volume of the box, using the correct number of significant figures in your answer.

SOLUTION

The volume of a box is determined by the product of its width, length and height. In reporting the product, we can show only as many significant figures as given in the dimension with the fewest significant figures, that for the height (two significant figures):

$$\begin{aligned}\text{Volume} &= \text{width} \times \text{length} \times \text{height}\\ &= (15.5\ \text{cm})\ (27.3\ \text{cm})\ (5.4\ \text{cm}) = 2285.01\ \text{cm}^3 \Rightarrow 2.3 \times 10^3\ \text{cm}^3\end{aligned}$$

When we use a calculator to do this calculation, the display shows 2285.01, which we must round off to two significant figures. Because the resulting number is 2300, it is best reported in exponential notation, 2.3×10^3, to indicate clearly two significant figures.

PRACTICE EXERCISE

It takes 10.5 s for a sprinter to run 100.00 m. Calculate the average speed of the sprinter in metres per second, and express the result to the correct number of significant figures.

Answer: $9.52\ \text{m s}^{-1}$ (three significant figures)

(See also Exercise 1.28.)

STRATEGIES IN CHEMISTRY

THE IMPORTANCE OF PRACTICE and ESTIMATING ANSWERS

If you listen to lectures and think to yourself, 'Yes, I understand that', be assured, that if you have done nothing else to reinforce what you heard, when questioned about the material the following day, you would not be able to explain it to someone. This is because listening to lectures or reading a textbook gives you *passive* knowledge. To become proficient in chemistry you must apply what you have learned into *active* knowledge by doing the sample exercises, the practice exercises and the problems at the end of the chapter. The more you practise doing problems, the more active your knowledge and the better your understanding of the principles of chemistry involved becomes. Remember, however, if you are stuck on a problem ask for help from your instructor, a tutor or a fellow student. Spending an inordinate amount of time on a single exercise that you cannot solve is rarely effective and leads to frustration.

When you are performing calculations to solve a problem, remember that a calculator is an important instrument that provides an answer quickly. However, the accuracy of the answer depends on the accuracy of the input. Very simply, if you were told to multiply 3.17×4.282, you should obtain the answer 13.6 (three significant figures). If you happened to obtain 135.74, you should immediately realise that the answer is wrong by a magnitude of 10 due to an incorrect input of data, because simply by rounding off the above numbers to 3 and 4 you would realise that the answer should be of the order of magnitude of 12. So by making a rough calculation using numbers that are rounded off in such a way that the arithmetic can be done without a calculator you will be able to check whether the answers to your calculations are reasonable. This requires some practice, but learning chemistry is all about practice as mentioned above.

When a calculation involves two or more steps and you write down answers for intermediate steps, retain at least one additional digit—past the number of significant figures—for the intermediate answers. This procedure ensures that small errors from rounding at each step do not combine to affect the final result. When using a calculator, you may enter the numbers one after another, rounding only the final answer. Accumulated rounding-off errors may account for small differences between results you obtain and answers given in the text for numerical problems.

SAMPLE EXERCISE 1.8 **Determining the number of significant figures in a calculated quantity**

A gas at 25 °C fills a container whose volume is 1.05×10^3 cm^3. The container plus gas have a mass of 837.6 g. The container, when emptied of all gas, has a mass of 836.2 g. What is the density of the gas at 25 °C?

SOLUTION

To calculate the density, we must know both the mass and the volume of the gas. The mass of the gas is just the difference in the masses of the full and empty container:

$$(837.6\ \text{g} - 836.2)\ \text{g} = 1.4\ \text{g}$$

In subtracting numbers, we determine the number of significant figures in our result by counting decimal places in each quantity. In this case each quantity has one decimal place. Thus the mass of the gas, 1.4 g, has one decimal place.

Using the volume given in the question, 1.05×10^3 cm^3, and the definition of density, we have

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{1.4\ \text{g}}{1.05 \times 10^3\ \text{cm}^3}$$

$$= 1.3 \times 10^{-3}\ \text{g cm}^{-3} = 0.0013\ \text{g cm}^{-3}$$

In dividing numbers, we determine the number of significant figures in our result by counting the number of significant figures in each quantity. There are two significant figures in our answer, corresponding to the smaller number of significant figures in the two numbers that form the ratio.

PRACTICE EXERCISE

To how many significant figures should the mass of the container be measured (with and without the gas) in Sample Exercise 1.8 in order for the density to be calculated to three significant figures?

Answer: Five; in order for the difference in the two masses to have three significant figures, there must be two decimal places in the masses of the filled and empty containers.

STRATEGIES IN CHEMISTRY

THE FEATURES OF THIS BOOK

To help you understand chemistry, this book includes features that help you organise your thoughts. At the beginning of each chapter, Key Concepts, which outline the chapter by section, will prepare you for the material in the chapter. At the end of each chapter, the Summary of Key Concepts, Key Skills, and Key Equations will help you remember what you have learned and prepare you for quizzes and exams.

The Concept Check features are placed in the text to test your understanding of what you have just read and the Figure It Out features are associated with artwork and ask you to interpret a concept visually. Sample Exercises, with worked-out solutions and answers, and Practice Exercises, which provide only the answer, test your problem-solving skills in chemistry.

At the end of each chapter is a series of exercises to allow you to practise your problem-solving skills further. The first few exercises, called Visualising Concepts, are meant to test how well you understand a concept without plugging a lot of numbers into a formula. The other exercises are divided into sections that reflect the order of the material in the chapter, with answers provided online. Additional Exercises appear after the regular exercises; the chapter sections that they cover are identified. Integrative Exercises, which start appearing in Chapter 3, are problems that require skills learned in previous chapters.

Throughout the book boxed essays highlight the importance of chemistry to our everyday lives. The My World of Chemistry boxes focus on biological and environmental and industrial aspects of chemistry. Strategies in Chemistry boxes, like this one, are meant to help you think about the material you are learning. Finally, boxes entitled A Closer Look provide in-depth coverage of a key chemical concept.

Many chemical databases are available, usually through your university or school. The *CRC Handbook of Chemistry and Physics* is the standard reference for many types of data and is available in libraries. The *Merck Index* is a standard reference for the properties of many small organic compounds, especially ones of biological interest. WebElements (www.webelements.com) is a good website for looking up the properties of the elements.

CHAPTER SUMMARY AND KEY TERMS

SECTION 1.1 Chemistry is the study of the composition, structure, properties and changes of **matter**. The composition of matter relates to the kind of elements it contains. The structure of matter relates to the ways the **atoms** of these elements are arranged. A **property** is any characteristic that gives a sample of matter its unique identity. A **molecule** is an entity composed of two or more atoms with the atoms attached to one another in a specific way.

SECTION 1.2 Matter exists in three physical states, **gas**, **liquid** and **solid**, which are known as the **states of matter**. There are two kinds of **pure substances**: **elements** and **compounds**. Each element has a single kind of atom and is represented by a **chemical symbol** consisting of one or two letters, with the first letter capitalised. Compounds are composed of two or more kinds of atoms joined in some specific fashion. The **law of constant composition**, also called the **law of definite proportion**, states that the elemental composition of a pure compound is always the same. Most matter consists of mixtures of pure substances. **Mixtures** can have variable composition and can be either homogeneous or heterogeneous. A homogeneous mixture is called a **solution**.

SECTION 1.3 Each form of matter (substance) has a unique set of **physical** and **chemical properties** that can be used to distinguish it from other forms of matter. **Physical changes**, such as **changes of state**, do not alter the composition of matter and are reversible, whereas **chemical changes** (**chemical reactions**) transform matter into chemically different substances. **Intensive properties** are independent of the amount of matter being examined and are used to identify substances. **Extensive properties** relate to the amount of substance present.

SECTION 1.4 Measurements in chemistry are made using the **metric system** using an internationally accepted system of units called **SI units**. SI units are based on the **metre** (m) and the **kilogram** (kg) as the basic units of length and mass. Although the SI temperature scale is the **Kelvin** (K) **scale**, the **Celsius** (°C) **scale** is most commonly used in chemistry. The SI unit for **volume** (V) is the cubic metre (m^3), but the cubic centimetre (cm^3) is commonly used in chemistry. **Density** (ρ) is an important quantity and is reported and equals mass divided by volume.

SECTION 1.5 All measured quantities are inexact to some extent. The **precision** of a measurement indicates how closely different measurements of a quantity agree with each other whereas **accuracy** of a measurement indicates how well a measurement agrees with the accepted or 'true' value. The number of **significant figures** (digits) in a measured quantity include one estimated digit, the last digit of the measurement. The significant figures indicate the degree of uncertainty in the measurement and must be reported according to a set of rules.

KEY SKILLS

- Distinguish between elements, compounds and mixtures. (Section 1.2)
- Memorise symbols of common elements and common prefixes for units. (Section 1.2)
- Memorise common SI units and metric prefixes. (Section 1.4)
- Use significant figures, scientific notation and SI units. (Section 1.5)

KEY EQUATIONS

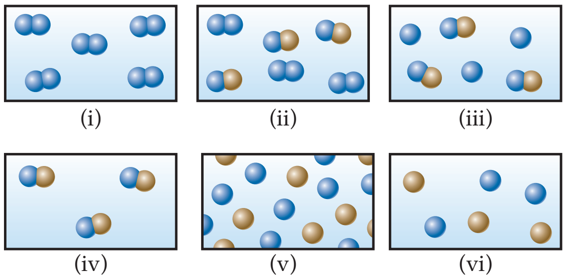
- Interconversion between Celsius (°C) and Kelvin (K) temperature scale $\quad K = °C + 273.15 \quad$ [1.1]
- Definition of density $\quad \text{Density} = \dfrac{\text{mass}}{\text{volume}};\ \rho = \dfrac{m}{V} \quad$ [1.2]

EXERCISES

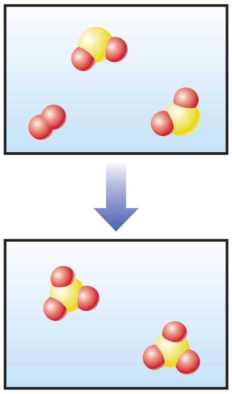
VISUALISING CONCEPTS

The exercises in this section are intended to probe your understanding of key concepts rather than your ability to utilise formulae and perform calculations.

1.1 Which of the following figures represents **(a)** a pure element, **(b)** a mixture of two elements, **(c)** a pure compound, **(d)** a mixture of an element and a compound? (More than one picture might fit each description.) [Section 1.2]



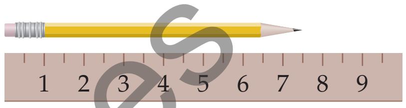
1.2 Does the following diagram represent a chemical or physical change? How do you know? [Section 1.3]



1.3 Identify each of the following as measurements of length, area, volume, mass, density, time or temperature: **(a)** 5 ns, **(b)** 5.5 kg m^{-3}, **(c)** 0.88 pm, **(d)** 540 km^2, **(e)** 173 K, **(f)** 2 mm^3, **(g)** 23 °C. [Section 1.4]

1.4 The following dartboards illustrate the types of errors often seen when one measurement is repeated several times. The bull's-eye represents the 'true value' and the darts represent the experimental measurements. Which board best represents each of the following scenarios: **(a)** measurements both accurate and precise, **(b)** measurements precise but inaccurate, **(c)** measurements imprecise but yield an accurate average? [Section 1.5]

1.5 What is the length of the pencil in the following figure if the scale reads in centimetres? How many significant figures are there in this measurement? [Section 1.5]



CLASSIFICATION AND PROPERTIES OF MATTER (Sections 1.2 and 1.3)

The following exercises are divided into sections that deal with specific topics in this chapter.

1.6 Classify each of the following as a pure substance or a mixture; if a mixture, indicate whether it is homogeneous or heterogeneous: **(a)** rice pudding, **(b)** seawater, **(c)** magnesium, **(d)** petrol.

1.7 Classify each of the following as a pure substance or a mixture; if a mixture, indicate whether it is homogeneous or heterogeneous: **(a)** air, **(b)** tomato juice, **(c)** iodine crystals, **(d)** sand.

1.8 Give the chemical symbols for the following elements: **(a)** sulfur, **(b)** potassium, **(c)** chlorine, **(d)** copper, **(e)** silicon, **(f)** nitrogen, **(g)** calcium, **(h)** helium.

1.9 Give the chemical symbol for each of the following elements: **(a)** carbon, **(b)** sodium, **(c)** fluorine, **(d)** iron, **(e)** phosphorus, **(f)** argon, **(g)** nickel, **(h)** silver.

1.10 Name the chemical elements represented by the following symbols: **(a)** Li, **(b)** Al, **(c)** Pb, **(d)** S, **(e)** Br, **(f)** Sn, **(g)** Cr, **(h)** Zn.

1.11 Name each of the following elements: **(a)** Co, **(b)** I, **(c)** Kr, **(d)** Hg, **(e)** As, **(f)** Ti, **(g)** K, **(h)** Ge.

1.12 A solid white substance A is heated strongly in the absence of air. It decomposes to form a new white substance B and a gas C. The gas has exactly the same properties as the product obtained when carbon is burned in an excess of oxygen. Based on these observations, can we determine whether solids A and B and the gas C are elements or compounds? Explain your conclusions for each substance.

1.13 In the process of attempting to characterise a substance, a chemist makes the following observations. The substance is a silvery white, lustrous metal. It melts at 649° C and boils at 1105 °C. Its density at 20 °C is 1.738 g cm^{-3}. The substance burns in air, producing an intense white light. It reacts with chlorine to give a brittle white solid. The substance can be pounded into thin sheets or drawn into wires. It is a good conductor of electricity. Which of these characteristics are physical properties and which are chemical properties?

1.14 Label each of the following as either a physical process or a chemical process: **(a)** corrosion of aluminium metal, **(b)** melting of ice, **(c)** pulverising an aspirin, **(d)** digesting a chocolate bar, **(e)** explosion of nitroglycerin.

1.15 Suggest a method of separating each of the following mixtures into two components: **(a)** sugar and sand, **(b)** iron and sulfur.

1.16 A beaker contains a clear, colourless liquid. If it is water, how could you determine whether it contained dissolved table salt? Do *not* taste it!

UNITS OF MEASUREMENT (Section 1.4)

1.17 Perform the following conversions: **(a)** 25.5 mg to g, **(b)** 4.0×10^{-10} m to nm, **(c)** 0.575 mm to µm.

1.18 Convert **(a)** 9.5×10^{-2} kg to g, **(b)** 0.0023 µm to nm, **(c)** 7.25×10^{-4} s to ms.

1.19 **(a)** A sample of carbon tetrachloride, a liquid once used in dry cleaning, has a mass of 39.73 g and a volume of 25.0 cm^3 at 25 °C. What is its density at this temperature? Will carbon tetrachloride float on water? (Materials that are less dense than water will float.) **(b)** The density of platinum is 21.45 g cm^{-3} at 20 °C. Calculate the mass of 75.00 cm^3 of platinum at this temperature. **(c)** The density of magnesium is 1.738 g cm^{-3} at 20 °C. What is the volume of 87.50 g of this metal at this temperature?

1.20 **(a)** A cube of osmium metal 1.500 cm on a side has a mass of 76.31 g at 25 °C. What is its density in g cm^{-3} at this temperature? **(b)** The density of titanium metal is 4.51 g cm^{-3} at 25 °C. What mass of titanium displaces 65.8 cm^3 of water at 25 °C? **(c)** The density of benzene at 15° C is 0.8787 g cm^{-3}. Calculate the mass of 0.1500 dm^3 of benzene at this temperature.

1.21 **(a)** To identify a liquid substance, a student determined its density. Using a graduated cylinder, she measured out a 45 cm^3 sample of the substance. She then measured the mass of the sample, finding that it weighed 38.5 g. She knew that the substance had to be either isopropyl alcohol (density 0.785 g cm^{-3}) or toluene (density 0.866 g cm^{-3}). What are the calculated density and the probable identity of the substance? **(b)** An experiment requires 45.0 g of ethylene glycol, a liquid whose density is 1.114 g cm^{-3}. Rather than weigh the sample on a balance, a chemist chooses to dispense the liquid using a graduated cylinder. What volume of the liquid should he use? **(c)** A cubic piece of metal measures 5.00 cm on each edge. If the metal is nickel, which has a density of 8.90 g cm^{-3}, what is the mass of the cube?

UNCERTAINTY IN MEASUREMENT (Section 1.5)

1.22 Indicate which of the following are exact numbers: **(a)** the mass of a paper clip, **(b)** the surface area of a coin, **(c)** the number of microseconds in a week, **(d)** the number of pages in this book.

1.23 Indicate which of the following are exact numbers: **(a)** the number of students in your chemistry class, **(b)** the temperature of the surface of the sun, **(c)** the mass of a postage stamp, **(d)** the number of millilitres in a cubic metre of water, **(e)** the average height of students in your class.

1.24 What is the number of significant figures in each of the following measured quantities? **(a)** 358 kg, **(b)** 0.054 s, **(c)** 6.3050 cm, **(d)** 0.0105 dm^3, **(e)** 7.0500×10^{-3} m^3.

1.25 Indicate the number of significant figures in each of the following measured quantities: **(a)** 3.7745 km, **(b)** 205 m^2, **(c)** 1.700 cm, **(d)** 350.0 K, **(e)** 307.080 g.

1.26 Round each of the following numbers to four significant figures, and express the result in standard exponential notation: **(a)** 102.53070, **(b)** 656980, **(c)** 0.008543210, **(d)** 0.000257870, **(e)** – 0.0357202.

1.27 **(a)** The diameter of Earth at the equator is 12 784.49 km. Round this number to three significant figures and express it in standard exponential notation. **(b)** The circumference of Earth through the poles is 40 008 km. Round this number to four significant figures and express it in standard exponential notation.

1.28 Compute the following and express the answers with the appropriate number of significant figures:
(a) $12.0550 + 9.05$
(b) $257.2 - 19.789$
(c) $(6.21 \times 10^3)\ (0.1050)$
(d) $0.0577/0.753$.

1.29 Compute the following and express the answer with the appropriate number of significant figures:
(a) $320.55 - (6104.5/2.3)$
(b) $[(285.3 \times 10^5) - (1.200 \times 10^3)] \times 2.8954$
(c) $(0.0045 \times 20\,000.0) + (2813 \times 12)$
(d) $863 \times [1255 - (3.45 \times 108)]$.

ADDITIONAL EXERCISES

The exercises in this section are not divided by category, although they are roughly in the order of the topics in the chapter.

1.30 What is meant by the terms 'composition' and 'structure' when referring to matter?

1.31 **(a)** Classify each of the following as a pure substance, a solution or a heterogeneous mixture: a gold coin, a cup of coffee, a wood plank. **(b)** What ambiguities are there in answering part **(a)** from the descriptions given?

1.32 **(a)** What is the difference between a hypothesis and a theory? **(b)** Explain the difference between a theory and a scientific law. Which addresses how matter behaves, and which addresses why it behaves that way?

1.33 The liquid substances mercury (density = 13.5 g cm^{-3}), water (1.00 g cm^{-3}) and cyclohexane (0.778 g cm^{-3}) do not form a solution when mixed, but separate in distinct layers. Sketch how the liquids would position themselves in a test tube.

1.34 **(a)** You are given a bottle that contains 4.59 cm^3 of a metallic solid. The total mass of the bottle and solid is 35.66 g. The empty bottle weighs 14.23 g. What is the density of the solid? **(b)** Mercury is traded by the 'flask', a unit that has a mass of 34.5 kg. What is the volume of a flask of mercury if the density of mercury is 13.5 g cm^{-3}? **(c)** A thief plans to steal a gold sphere with a radius of 28.9 cm from a museum. If the gold has a density of 19.3 g cm^{-3} what is the mass of the sphere? (The volume of a sphere is $V = (4/3)\pi r^3$.) Is he likely to be able to walk off with it unassisted?

1.35 Car batteries contain sulfuric acid, which is commonly referred to as 'battery acid'. Calculate the number of grams of sulfuric acid in 0.500 dm^3 of battery acid if the solution has a density of 1.28 g cm^{-3} and is 38.1% sulfuric acid by mass.

1.36 Gold is alloyed (mixed) with other metals to increase its hardness in making jewellery. **(a)** Consider a piece of gold jewellery that weighs 9.85 g and has a volume of 0.0675 cm^3. The jewellery contains only gold and silver, which have densities of 19.3 g cm^{-3} and 10.5 g cm^{-3}, respectively. Assuming that the total volume of the jewellery is the sum of the volumes of the gold and silver that it contains, calculate the percentage of gold (by mass) in the jewellery. **(b)** The relative amount of gold in an alloy is commonly expressed in units of carats. Pure gold is 24 carat, and the percentage of gold in an alloy is given as a percentage of this value. For example, an alloy that is 50% gold is 12 carat. State the purity of the gold jewellery in carats.

1.37 Suppose you are given a sample of a homogeneous liquid. What would you do to determine whether it is a solution or a pure substance?

PHOTO/ART CREDITS

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