Orange juice, Coca-Cola, and pancake syrup are mixtures that contain sugar. The amount of sugar varies from one to another because mixtures can have variable compositions.

but it supports combustion, whereas carbon dioxide is used to extinguish fires. Thus, the properties of carbon dioxide are much different from those of both carbon and oxygen. When carbon and oxygen combine to form carbon dioxide, the properties of carbon and oxygen disappear and in their place we find the unique properties of the compound carbon dioxide.

Mixtures can have variable compositions

Elements and compounds are examples of **pure substances.**[6] The composition of a pure substance is always the same, regardless of its source. All samples of table salt, for example, contain sodium and chlorine combined in the same proportions by mass. Similarly, all samples of water contain the same proportions by mass of hydrogen and oxygen. Pure substances are rare, however. Usually we encounter mixtures of compounds or elements. Unlike elements and compounds, **mixtures** *can have variable compositions.* Carbon dioxide and water are examples of compounds because each consists of elements chemically combined in definite proportions. But, when we dissolve carbon dioxide in water to give a carbonated beverage such as seltzer, we can vary the amounts of the two compounds. As a result, we can vary the amount of "fizz" in the beverage.

Mixtures can be either homogeneous or heterogeneous. A **homogeneous mixture** *has the same properties throughout the sample.* An example is a thoroughly stirred mixture of sugar in water. We call such a homogeneous mixture a **solution.** Solutions need not be liquids, just homogeneous. Brass, for example, is a solid solution of copper and zinc, and clean air is a gaseous solution of oxygen, nitrogen, and a number of other gases.

A **heterogeneous mixture** *consists of two or more regions, called* **phases,** *that differ in properties.* A mixture of olive oil and vinegar in a salad dressing, for example, is a two-phase mixture in which the oil floats on the vinegar as a separate layer (Figure 1.7). The phases in a mixture don't have to be chemically different substances like oil and vinegar, however. A mixture of ice and liquid water is a two-phase heterogeneous mixture in which the phases have the same chemical composition but occur in different physical states.

An important way that mixtures differ from compounds is in the changes that occur when they form. Consider, for example, the elements iron and sulfur, which are pictured in powdered form in Figure 1.8*a.* We can make a mixture simply by dumping them together and stirring them. In the mixture (Figure 1.8*b*), both elements retain their original properties. The process we use to create this mixture

FIGURE 1.7 ***A heterogeneous mixture.*** The salad dressing shown here contains vinegar and vegetable oil (plus assorted other flavorings). Vinegar and oil do not dissolve in each other but form two layers. The mixture is heterogeneous because each of the separate phases (oil, vinegar, and other solids) has its own set of properties that differ from the properties of the other phases.

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

(*a*) (*b*)

FIGURE 1.8 ***Formation of a mixture of iron and sulfur.*** (*a*) Samples of powdered sulfur and powdered iron. (*b*) A mixture of sulfur and iron is made by stirring the two powders together.

FIGURE 1.9 ***Formation of a mixture is a physical change.*** Here we see that forming the mixture has not changed the iron and sulfur into a compound of these two elements. The mixture can be separated by pulling the iron out with a magnet.

involves a *physical change,* rather than a chemical change, because no new chemical substances form. To separate the mixture, we could similarly use just physical changes. For example, we could remove the iron by stirring the mixture with a magnet—a physical operation. The iron powder sticks to the magnet as we pull it out, leaving the sulfur behind (Figure 1.9). The mixture also could be separated by treating it with a liquid called carbon disulfide, which is able to dissolve the sulfur but not the iron. Filtering the sulfur solution from the solid iron, followed by evaporation of the liquid carbon disulfide from the sulfur solution, gives the original components, iron and sulfur, separated from each other.

As we noted earlier, formation of a compound from its elements involves a chemical reaction. Iron and sulfur, for example, combine to form a compound often called "fool's gold" because of its appearance (Figure 1.5, page 8). In this compound the elements no longer have the same properties they had before they were combined, and they cannot be separated by physical means. Just as the formation of a compound involves a chemical reaction, so also does its decomposition. The decomposition of fool's gold into iron and sulfur is therefore a chemical change.

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

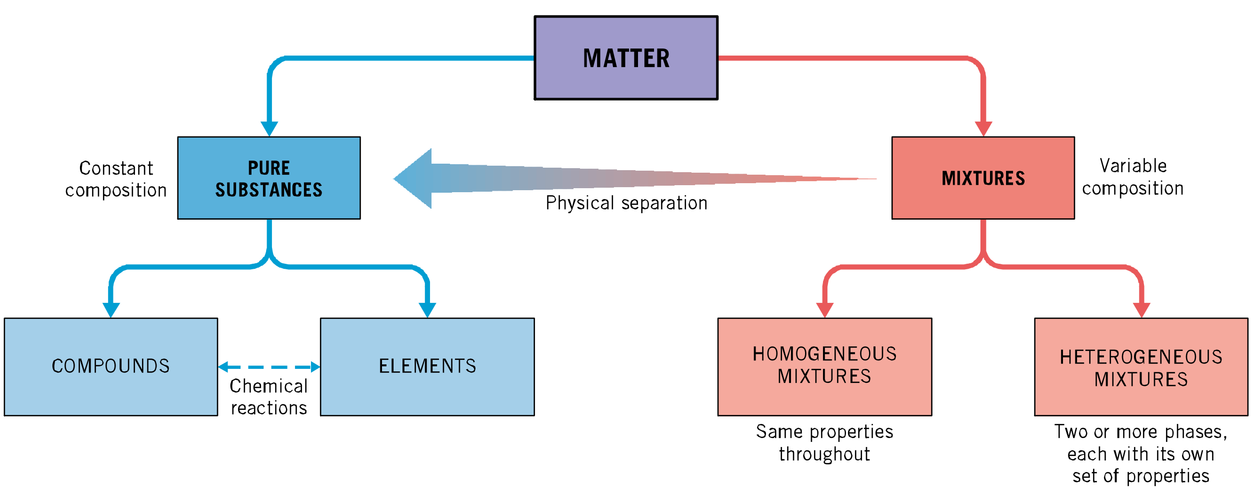

FIGURE 1.10 ***Classification of matter.***

1.5 Atoms of an element have properties in common

In our discussion of elements in the preceding section, no reference was made to the atomic nature of matter. In fact, the distinction between elements and compounds had been made even before the atomic theory of matter was formulated. In this section we will examine how the atomic theory began and take a closer look at elements in terms of our modern view of atomic structure.

Dalton's atomic theory explained chemical laws

In modern science, we have come to take for granted the existence of atoms and molecules. In fact, we've already used the atomic theory to explain some of the properties of materials. However, scientific evidence for the existence of atoms is relatively recent, and chemistry did not progress very far until that evidence was found. Therefore, let's take a brief look at how the atomic theory evolved.

The concept of atoms began nearly 2500 years ago when certain Greek philosophers expressed the belief that matter is ultimately composed of tiny indivisible particles, and it is from the Greek word *atomos,* meaning "not cut," that the word *atom* is derived. The philosophers' conclusions, however, were not supported by any evidence; they were derived simply from philosophical reasoning.

Laws of chemical combination evolved from experimental observations

The concept of atoms remained a philosophical belief, having limited scientific usefulness, until the discovery of two quantitative laws of chemical combination—the *law of conservation of mass* and the *law of definite proportions.* The evidence that led to the discovery of these laws came from the experimental observations of many scientists in the eighteenth and early nineteenth centuries.

Law of Conservation of Mass
No detectable gain or loss of mass occurs in chemical reactions. Mass is *conserved.*

Law of Definite Proportions
In a given chemical compound, the elements are always combined in the same proportions by mass.

TOOLS

Laws of definite proportions and conservation of mass

Notice that both of these laws refer to the masses of substances. Earlier, in our definition of matter, we noted that mass is a measure of the amount of matter in an object. In the sciences, we measure mass in units of **grams** (symbol, **g**). In Chapter 3 we will discuss units of measurement, including those of mass, in greater detail, but for now grams are sufficient.

In the United States, we commonly measure mass in units of ounces or pounds. One pound equals 453.6 g. A penny has a mass of about 3 g.

The **law of conservation of mass** means that if a chemical reaction takes place in a sealed vessel that permits no matter to enter or escape, the mass of the vessel and its contents after the reaction will be identical to its mass before. Although this may seem quite obvious to us now, it wasn't quite so clear in the early history of modern chemistry. Not until scientists made sure that *all* substances, including any that were gases, were included when masses were measured could the law of conservation of mass be truly tested.

We actually used the **law of definite proportions** on page 10 when we defined a compound as a substance in which two or more elements are chemically combined in a *definite fixed proportion by mass.* Thus, if we decompose samples of water (a compound) into the elements oxygen and hydrogen, we always find that the ratio of oxygen to hydrogen, *by mass,* is 8 to 1. In other words, the mass of oxygen obtained is always eight times the mass of hydrogen.

9 g water → 1 g hydrogen, 8 g oxygen

18 g water → 2 g hydrogen, 16 g oxygen

In any sample of water, the mass of oxygen is always eight times the mass of hydrogen.

Similarly, if we form water from oxygen and hydrogen, the mass of oxygen consumed will always be eight times the mass of hydrogen that reacts. This is true even if there's a large excess of one of them. For instance, if 100 g of oxygen are mixed with 1 g of hydrogen and the reaction to form water is initiated, all the hydrogen would react but only 8 g of oxygen would be consumed; there would be 92 g of oxygen left over. No matter how we try, we can't alter the chemical composition of the water formed in the reaction.

100 g oxygen 1 g hydrogen	$\xrightarrow{\text{chemical reaction}}$	9 g water 92 g oxygen
mixture before reaction	occurs consuming 8 g oxygen 1 g hydrogen	mixture after reaction

Let's look at a sample calculation that shows how we might use the law of definite proportions.

EXAMPLE 1.1

Applying the Law of Definite Proportions

The element molybdenum (Mo) combines with sulfur (S) to form a compound commonly called molybdenum disulfide that is useful as a dry lubricant, similar to graphite. It is also used in specialized lithium batteries. A sample of this compound contains 1.50 g of Mo for each 1.00 g of S. If a different sample of the compound contains 2.50 g of S, how many grams of Mo does it contain?

A Word about Problem Solving

This is the first of many encounters you will have with solving problems in chemistry. Helping you learn how to approach and solve problems is one of the major goals of this textbook. We view problem solving as a three-step process. The first is figuring out what has to be done to solve the problem, which is the function of the *Analysis* step described below. The second is actually performing whatever is required to obtain the answer (the *Solution* step). And finally, we examine the answer to determine whether it seems to be *reasonable.* For more information on the aids that are available to assist you in problem solving, we recommend that you read the "To The Student" section at the beginning of the book.

ANALYSIS: Much of the effort in solving a chemistry problem is devoted to determining which concepts have to be applied. We view these concepts as *tools,* each with its specific uses when applied to problem solving. Our goal in this Analysis step is to identify which tools we need and how we will apply them.

We begin by examining the problem and asking a question: What have we learned that relates the masses of elements in two samples of the same compound? The law of definite proportions states that the proportions of the elements by mass must be the same in both samples, so the law of definite proportions is the tool we need to apply. The law tells us that the ratio of grams of S to grams of Mo must be the same in both samples. To solve the problem, then, we will set up the mass ratios for the two samples. In the ratio for the second sample, the mass of molybdenum will be an unknown quantity. We'll equate the two ratios and solve for the unknown quantity.

SOLUTION: Now that we've determined what we need to do to solve the problem, the rest is pretty easy. The first sample has a Mo to S mass ratio of

$$\frac{1.50 \text{ g Mo}}{1.00 \text{ g S}}$$

In the second sample, we know the mass of S (2.50 g) and we want to find the mass of Mo, so let's call this quantity x. The mass ratio of Mo to S in the second sample is therefore

$$\frac{x}{2.50 \text{ g S}}$$

Now we equate them, because the two ratios must be equal.

$$\frac{x}{2.50 \text{ g S}} = \frac{1.50 \text{ g Mo}}{1.00 \text{ g S}}$$

Solving for x gives

$$x = 2.50 \text{ g S} \times \frac{1.50 \text{ g Mo}}{1.00 \text{ g S}} = 3.75 \text{ g Mo}$$

Is the Answer Reasonable?

To avoid errors, it's always wise to do a rough check of the answer. Usually, some simple reasoning is all we need to see if the answer is "in the right ball park." This is how we might do such a check here: Notice that the amount of sulfur in the second sample is more than twice the amount in the first sample. Therefore, we should expect the amount of Mo in the second sample to be somewhat more than twice what it is in the first. The answer we obtained, 3.75 g Mo, is more than twice 1.50 g Mo, so our answer seems to be reasonable.

PRACTICE EXERCISE 2: Cadmium sulfide is a yellow compound that is used as a pigment in artist's oil colors. A sample of this compound is composed of 1.25 g of cadmium and 0.357 g of sulfur. If a second sample of the same compound contains 3.50 g of sulfur, how many grams of cadmium does it contain?

The atomic theory was proposed by John Dalton

The laws of conservation of mass and definite proportions served as the *experimental foundation* for the atomic theory. They prompted the question, "What must be true about the nature of matter, given the truth of these laws?" In other words, what is matter made of?

At the beginning of the nineteenth century, John Dalton (1766–1844), an English scientist, used the Greek concept of atoms to make sense out of the laws of conservation of mass and definite proportions. Dalton reasoned that if atoms really exist, they must have certain properties to account for these laws. He described such properties, and the list constitutes what we now call **Dalton's atomic theory.**

Dalton's Atomic Theory

1. *Matter consists of tiny particles called atoms.*
2. *Atoms are indestructible. In chemical reactions, the atoms rearrange but they do not themselves break apart.*
3. *In any sample of a pure element, all the atoms are identical in mass and other properties.*
4. *The atoms of different elements differ in mass and other properties.*
5. *When atoms of different elements combine to form compounds, new and more complex particles form. However, in a given compound the constituent atoms are always present in the same fixed* ***numerical*** *ratio.*

Dalton's theory easily explained the law of conservation of mass. According to the theory, a chemical reaction is simply a reordering of atoms from one combination to another. If no atoms are gained or lost and if the masses of the atoms can't change, then the mass after the reaction must be the same as the mass before. This explanation of the law of conservation of mass works so well that it serves as the reason for balancing chemical equations, which we will discuss in the next chapter.

The law of definite proportions is also easy to explain. According to the theory, a given compound always has atoms of the same elements in the same numerical ratio. Suppose, for example, that two elements, A and B, combine to form a compound in which the number of atoms of A equals the number of atoms of B (i.e., the *atom ratio* is 1 to 1). If the mass of a B atom is twice that of an A atom, then every time we encounter a sample of this compound, the mass ratio (A to B) would be 1 to 2. This same mass ratio would exist regardless of the size of the sample, so in samples of this compound the elements A and B are always present in the same proportion by mass.

The atomic theory led to the discovery of the law of multiple proportions

Strong support for Dalton's theory came when Dalton and other scientists studied elements that are able to combine to give two (or more) compounds. For example, sulfur and oxygen form two different compounds, which we call sulfur dioxide and sulfur trioxide. If we decompose a 2.00 g sample of sulfur dioxide, we find it contains 1.00 g of S and 1.00 g of O. If we decompose a 2.50 g sample of sulfur trioxide, we find it also contains 1.00 g of S, but this time the mass of O is 1.50 g. This is summarized in the following table.

Compound	Sample Size	Mass of Sulfur	Mass of Oxygen
Sulfur dioxide	2.00 g	1.00 g	1.00 g
Sulfur trioxide	2.50 g	1.00 g	1.50 g

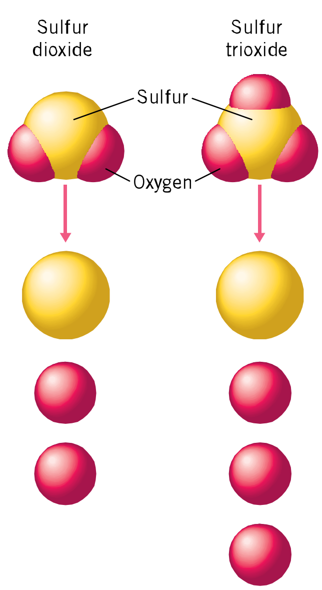

FIGURE 1.11 ***Oxygen compounds of sulfur demonstrate the law of multiple proportions.*** Illustrated here are molecules of sulfur trioxide and sulfur dioxide. Each has one sulfur atom, and therefore the same mass of sulfur. The oxygen ratio is 3 to 2, both by atoms and by mass.

First, notice that sample sizes aren't the same; they were chosen so that each has the *same mass of sulfur.* Second, the ratio of the masses of oxygen in the two samples is one of small whole numbers.

$$\frac{\text{mass of oxygen in sulfur trioxide}}{\text{mass of oxygen in sulfur dioxide}} = \frac{1.50\text{ g}}{1.00\text{ g}} = \frac{3}{2}$$

Similar observations are made when we study other elements that form more than one compound with each other, and these observations form the basis of the **law of multiple proportions.**

Law of Multiple Proportions
Whenever two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers.

Dalton's theory explains the law of multiple proportions in a very simple way. Suppose a molecule of sulfur trioxide contains one sulfur and three oxygen atoms, and a molecule of sulfur dioxide contains one sulfur and two oxygen atoms (Figure 1.11). If we had just one molecule of each, then our samples each would have one sulfur atom and therefore the same mass of sulfur. Then, comparing the oxygen atoms, we find they are in a numerical ratio of 3 to 2. But because oxygen atoms all have the same mass, the mass ratio must also be 3 to 2.

The law of multiple proportions was not known before Dalton presented his theory, and its discovery demonstrates the scientific method in action. Experimental data suggested to Dalton the existence of atoms, and the atomic theory suggested the relationships that we now call the law of multiple proportions. When found by experiment, the existence of the law of multiple proportions added great support to the atomic theory. In fact, for many years it was one of the strongest arguments in favor of the existence of atoms.

Modern experimental evidence exists for atoms

Atoms are so incredibly tiny that even the most powerful optical microscopes are unable to detect them. In recent times, though, scientists have developed very sensitive instruments that are able to map the surfaces of solids with remarkable resolution. One such instrument is called a **scanning tunneling microscope.** It was invented in the early 1980s by Gerd Binnig and Heinrich Rohrer and earned them the 1986 Nobel prize in physics. With this instrument, the tip of a sharp metal probe is brought very close to an electrically conducting surface and an electric current bridging the gap is begun. The flow of current is extremely sensitive to the distance between the tip of the probe and the sample. As the tip is moved across the surface, the height of the tip is continually adjusted to keep the current flow constant. By accurately recording the height fluctuations of the tip, a map of the hills and valleys on the surface is obtained. The data are processed using a computer to reveal images such as that shown in Figure 1.12.

Relative atomic masses of elements can be found

One of the most useful concepts to come from Dalton's atomic theory is that atoms of an element have a constant, characteristic **atomic mass** (or **atomic weight**). This concept opened the door to the determination of chemical formulas and ultimately to one of the most useful devices chemists have for organizing chemical information, the periodic table of the elements. But how can the masses of atoms be measured?

Individual atoms are much too small to weigh in the traditional manner. However, the *relative masses* of the atoms of elements can be determined *provided we know the ratio in which the atoms occur in a compound.* Let's look at an example to see how this could work.

Hydrogen (H) combines with the element fluorine (F) to form the compound hydrogen fluoride. Each molecule of this compound contains one atom of hydrogen and one atom of fluorine, which means that in *any* sample of this substance the fluorine-to-hydrogen *atom ratio* is always 1 to 1. It is also found that when a sample of hydrogen fluoride is decomposed, the mass of fluorine obtained is always 19.0

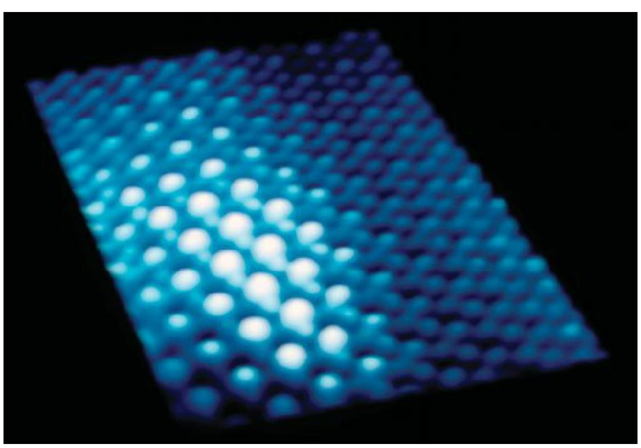

FIGURE 1.12 ***Individual atoms can be imaged using a scanning tunneling microscope.*** This STM micrograph reveals the pattern of individual atoms of palladium deposited on a graphite surface. Palladium is a silver-white metal used in alloys such as white gold and dental crowns.

times larger than the mass of hydrogen, so the fluorine-to-hydrogen *mass ratio* is always 19.0 to 1.00.

F-to-H atom ratio: 1 to 1

F-to-H mass ratio: 19.0 to 1.00

How could a 1-to-1 atom ratio give a 19.0-to-1.00 mass ratio? *Only if each fluorine atom is 19.0 times heavier than each H atom.*

Notice that even though we haven't found the actual masses of F and H atoms, we do now know how their masses compare (i.e., we know their *relative masses*). Similar procedures with other elements in other compounds are able to establish relative mass relationships among the other elements as well. What we need next is a way to place all these masses on the same mass scale, but before we study this, let's take a closer look at one of Dalton's hypotheses.

Isotopes are atoms of the same element with different masses

The cornerstone of Dalton's theory was the idea that all of the atoms of an element have identical masses. Actually, most elements occur in nature as uniform mixtures of two or more kinds of atoms that have slightly different masses, which we call **isotopes.** An iron nail, for example, is made up of a mixture of four iron isotopes, and the chlorine in salt is a mixture of two chlorine isotopes.

The existence of isotopes did not affect the development of Dalton's theory for two reasons. First, all the isotopes of a given element have virtually identical *chemical* properties—all give the same kinds of chemical reactions. Second, the relative proportions of its different isotopes are essentially constant, regardless of where on earth or in the atmosphere the element is found. As a result, every sample of a particular element has the same isotopic composition, and the *average* mass per atom is the same from sample to sample. Therefore, in the laboratory elements *behave* as though their atoms have masses equal to the average.

Even the smallest laboratory sample of an element has so many atoms that the relative proportions of the isotopes is constant.

Carbon-12 is the standard on the atomic mass scale

To establish a uniform mass scale for atoms it is necessary to select a standard against which the relative masses can be compared. Currently, the agreed-upon reference is the most abundant isotope of carbon, called carbon-12 and symbolized ^{12}C. One atom of this isotope is assigned *exactly* 12 units of mass, which are called **atomic mass units.** Some prefer to use the symbol **amu** for the atomic mass unit. The internationally accepted symbol is **u,** which is the symbol we will use throughout the rest of the book. By assigning 12 u to the mass of one atom of ^{12}C, the size of the atomic mass unit is established to be $\frac{1}{12}$ of the mass of a single carbon-12 atom:

The atomic mass unit is sometimes called a dalton.

1 u = 1 dalton

$$1 \text{ atom of } ^{12}\text{C has a mass of 12 u (exactly)}$$
$$1 \text{ u equals } \tfrac{1}{12} \text{ the mass of 1 atom of } ^{12}\text{C (exactly)}$$

In modern terms, the atomic mass of an element is the average mass of the element's atoms (as they occur in nature) relative to an atom of carbon-12, which is assigned a mass of 12 units. Thus, if an average atom of an element has a mass twice that of a ^{12}C atom, its atomic mass would be 24 u.

The definition of the size of the atomic mass unit is really quite arbitrary. It could just as easily have been selected to be $\frac{1}{24}$ of the mass of a carbon atom, or $\frac{1}{10}$ of the mass of an iron atom, or any other value. Why $\frac{1}{12}$ of the mass of a ^{12}C atom? First, carbon is a very common element, available to any scientist. Second, and most important, by choosing the atomic mass unit of this size, the atomic masses of nearly all the other elements are almost whole numbers, with the lightest atom (hydrogen) having a mass of approximately 1 u.

Chemists generally work with whatever *mixture* of isotopes comes with a given element as it occurs naturally. Because the composition of this isotopic mixture is very nearly constant regardless of the source of the element, we can speak of an *average atom* of the element—average in terms of mass. For example, naturally occurring hydrogen is a mixture of two isotopes in the relative proportions given in the margin. The "average atom" of the element hydrogen, as it occurs in nature, has a mass that is 0.083992 times that of a ^{12}C atom. Since 0.083992×12.000 u = 1.0079 u, the average atomic mass of hydrogen is 1.0079 u. Notice that this average value is just a little larger than the atomic mass of ^{1}H because naturally occurring hydrogen also contains a little ^{2}H.

Hydrogen Isotope	Mass	Percentage Abundance
^{1}H	1.007825 u	99.985
^{2}H	2.0140 u	0.015

Average atomic masses can be calculated from isotopic abundances

Originally, the relative atomic masses of the elements were determined in a way similar to that described for hydrogen and fluorine in our earlier discussion. A sample of a compound was analyzed and from the formula of the substance the relative atomic masses were calculated. These were then adjusted to place them on the unified atomic mass scale. In modern times, methods have been developed to measure very precisely both the relative abundances of the isotopes of the elements and their atomic masses. This kind of information has permitted the calculation of more precise values of the average atomic masses, which are found in the table on the inside front cover of the book. Example 1.2 illustrates how this calculation is done.

EXAMPLE 1.2

Calculating Average Atomic Masses from Isotopic Abundances

Naturally occurring chlorine is a mixture of two isotopes. In every sample of this element, 75.77% of the atoms are ^{35}Cl and 24.23% are atoms of ^{37}Cl. The accurately measured atomic mass of ^{35}Cl is 34.9689 u and that of ^{37}Cl is 36.9659 u. From these data, calculate the average atomic mass of chlorine.

ANALYSIS: This problem is a little like calculating the average height of a group of people. Let's see how we might do that and then apply the same principles to the problem at hand.

Suppose 10% of a group of people are 4 ft tall and the other 90% are 6 ft tall. The average height of the group is *not* 5 ft (an average of 4 ft and 6 ft). To calculate the average, we have to weight it according to the proportions of 4- and 6-footers. If there were 100 people in the group, 10 of them (which is 10% of 100) would be 4 ft tall and the other 90 would be 6 ft tall. To compute the average, we add all the heights of all the people and then divide by the total number. Let's see how this works out when we set up the math.

$$\frac{(10 \times 4\text{ ft}) + (90 \times 6\text{ ft})}{100} = \frac{(10 \times 4\text{ ft})}{100} + \frac{(90 \times 6\text{ ft})}{100}$$
$$= (0.1 \times 4\text{ ft}) + (0.9 \times 6\text{ ft})$$
$$= 0.4\text{ ft} + 5.4\text{ ft} = 5.8\text{ ft}$$

Notice that in this calculation, multiplying 4 ft by 0.1 is the same as taking 10% of 4 ft. Similarly, multiplying 6 ft by 0.9 is the same as taking 90% of 6 ft. Thus, we could say that the "average person" in this group (who really doesn't exist) has a height that's 10% of 4 ft plus 90% of 6 ft. With this in mind, let's see how we can apply it to calculating the average atomic mass.

In a sample containing many atoms of chlorine, 75.77% of the mass is contributed by atoms of ^{35}Cl and 24.23% is contributed by atoms of ^{37}Cl. This means that when we calculate the mass of the "average atom," we have to weight it according to both the masses of the isotopes and their relative abundances. As in the discussion above, to do this it is convenient to imagine an "average atom" to be composed of 75.77% of ^{35}Cl and 24.23% of ^{37}Cl. (Keep in mind, of course, that such an atom doesn't really exist. This is just a simple way to see how we can calculate the average atomic mass of this element.)

SOLUTION: We will calculate 75.77% of the mass of an atom of ^{35}Cl, which is the contribution of this isotope to the "average atom." Then we will calculate 24.23% of the mass of an atom of ^{37}Cl, which is the contribution of this isotope. Adding these contributions gives the total mass of the "average atom."

decimal form of 75.77%

$$0.7577 \times 34.9689 \text{ u} = 26.50 \text{ u} \qquad (\text{for } ^{35}\text{Cl})$$

decimal form of 24.23%

$$0.2423 \times 36.9659 \text{ u} = 8.957 \text{ u} \qquad (\text{for } ^{37}\text{Cl})$$

$$\text{total mass of average atom} = 35.46 \text{ u (rounded)}$$

The average atomic mass of chlorine is therefore 35.46 u.

Is the Answer Reasonable?

Once again, the final step is a check to see if the answer makes sense. Here is how we might do such a check: First, from the masses of the isotopes, we know the average atomic mass is somewhere between approximately 35 and 37. If the abundances of the two isotopes were equal, the average would be nearly 36. But there is more ^{35}Cl than ^{37}Cl, so a value somewhere between 35 and 36 seems reasonable; therefore, we can feel pretty confident our answer is correct.

PRACTICE EXERCISE 3: Aluminum atoms have a mass that is 2.24845 times that of an atom of ^{12}C. What is the atomic mass of aluminum?

PRACTICE EXERCISE 4: How much heavier than an atom of ^{12}C is the average atom of naturally occurring copper? Refer to the table inside the front cover of the book for the necessary data.

PRACTICE EXERCISE 5: Naturally occurring boron is composed of 19.8% of ^{10}B and 80.2% of ^{11}B. Atoms of ^{10}B have a mass of 10.0129 u and those of ^{11}B have a mass of 11.0093 u. Calculate the average atomic mass of boron.

1.6 Atoms are composed of subatomic particles

The earliest theories about atoms imagined them to be indestructible and totally unable to be broken into smaller pieces. However, as you probably know, atoms are not quite as indestructible as Dalton had thought. During the late 1800s and early 1900s, experiments were performed that demonstrated that atoms are composed instead of simpler **subatomic particles.** (For some of the details about these experiments, see Facets of Chemistry 1.1 and 1.2.) From this work the current theoretical model of atomic structure evolved. We will examine it in general terms in this chapter. A more detailed discussion of atomic structure will follow in Chapter 8.

Protons, neutrons, and electrons are subatomic particles

Experiments have shown that atoms are composed of three principal kinds of subatomic particles: **protons, neutrons,** and **electrons.** Experiments also revealed that at the center of an atom there exists a very tiny, extremely dense core called the **nucleus,** which is where an atom's protons and neutrons are found. Because they are found in nuclei, protons and neutrons are sometimes called **nucleons.** The elec-