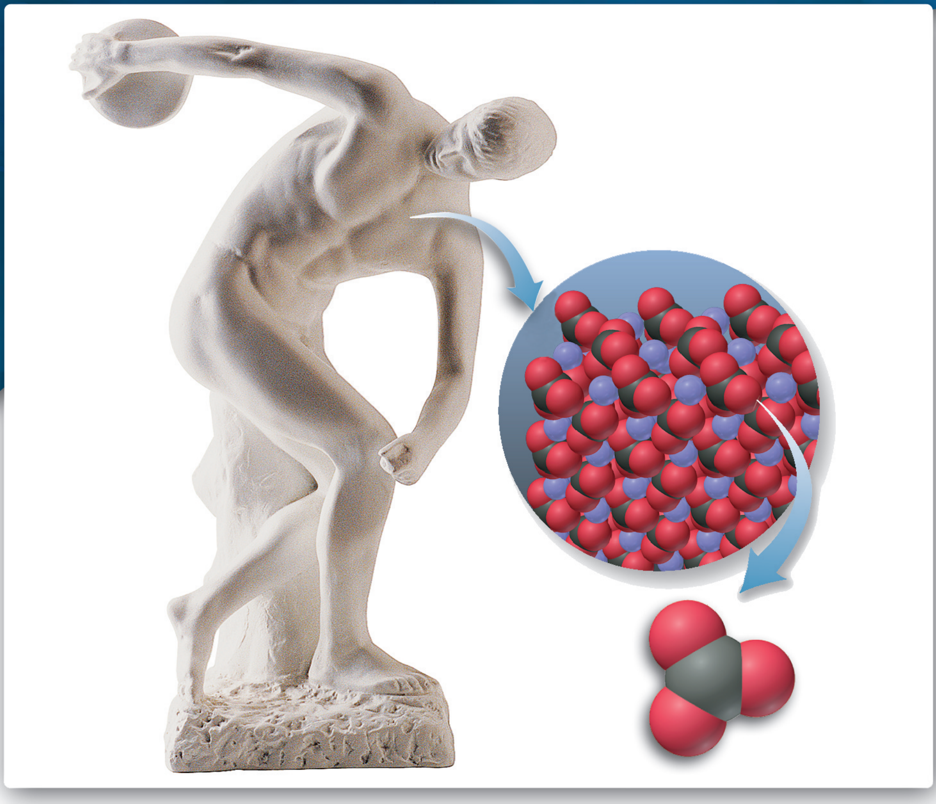
2



Great Art on the Molecular Scale. Any sample of matter has components, which in turn have smaller components, and so on down to the subatomic scale. This magnificent ancient Roman statue of a discus thrower is composed largely of marble, which consists of calcium and carbonate ions attracted to each other in a regular array. Each carbonate ion consists of carbon *(black)* and oxygen *(red)* atoms bonded to each other. In this chapter, you'll learn what matter is made of.

The Components of Matter

It may seem surprising, but questioning what things are made of is as common today as it was among the philosophers of ancient Greece, even though we approach the question very differently. They believed everything was made of one or, at most, a few elemental substances (elements). Some believed the elemental substance was water because rivers and oceans extend everywhere. Others thought it was air, which was "thinned" into fire or "thickened" into clouds, rain, and rock. Still others believed there were four elements—fire, air, water, and earth—whose properties accounted for hotness, wetness, sweetness, and all other characteristics of things.

Democritus (c. 460–370 BC), the father of atomism, took a different approach. He focused on the ultimate components of *all* substances, and his reasoning went something like this: if you cut a piece of, say, copper smaller and smaller, you must eventually reach a particle of copper so small that it can no longer be cut. Therefore, matter is ultimately composed of indivisible particles, with nothing between them but empty space. He called the particles *atoms* (Greek *atomos,* "uncuttable") and proclaimed: "According to convention, there is a sweet and a bitter, a hot and a cold, and according to convention, there is order. In truth, there are atoms and a void." However, Aristotle (384–322 BC), who elaborated the idea of four elements, held that it was impossible for "nothing" to exist, and his influence was so great that the concept of atoms was suppressed for 2000 years.

Finally, in the 17th century, the great English scientist Robert Boyle argued that an element is composed of "simple Bodies, not made of any other Bodies, of which all mixed Bodies are compounded, and into which they are ultimately resolved," a description that is remarkably similar to today's idea of an element, in which the "simple Bodies" are atoms. Boyle's hypothesis began the wonderful process of discovery, debate, and rediscovery that marks scientific inquiry, as exemplified by Lavoisier's work. Further studies in the 18th century gave rise to laws concerning the relative masses of substances that react with each other. Then, at the beginning of the 19th century, John Dalton proposed an atomic model that explained these mass laws and soon led to rapid progress in chemistry. By that century's close, however, further observation exposed the need to revise Dalton's model. A burst of creativity in the early 20th century gave rise to a picture of the atom with a complex internal structure, which led to our current model.

IN THIS CHAPTER . . . We compare the properties and composition of the three types of matter—elements, compounds, and mixtures—on the macroscopic and atomic scales. We examine the mass laws and Dalton's theory to explain them and then cover key experiments that led to our current model of the atom. Atomic structure is described, and then we see how elements are organized and classified in the periodic table. We discuss the two ways elements combine to form compounds, and learn how to derive compound names, formulas, and masses. Finally, we see how mixtures are classified and separated.

Concepts & Skills to Review before you study this chapter

- physical and chemical change (Section 1.1)
- states of matter (Section 1.1)
- attraction and repulsion between charged particles (Section 1.1)
- meaning of a scientific model (Section 1.3)
- SI units and conversion factors (Section 1.5)
- significant figures in calculations (Section 1.6)

2.1 ELEMENTS, COMPOUNDS, AND MIXTURES: AN ATOMIC OVERVIEW

Matter can be classified into three types based on its composition—elements, compounds, and mixtures. An **element** is the simplest type of matter with unique physical and chemical properties. *An element consists of only one kind of atom.* Therefore, it cannot be broken down into a simpler type of matter by any physical or chemical methods. An element is one kind of **substance,** matter whose composition is fixed. Each element has a name, such as silicon, oxygen, or copper. A sample of silicon contains only silicon atoms. A key point to remember is that the *macroscopic* properties of a piece of silicon, such as color, density, and combustibility, are different from those of a piece of copper because silicon atoms are different from copper atoms; in other words, *each element is unique because the properties of its atoms are unique.*

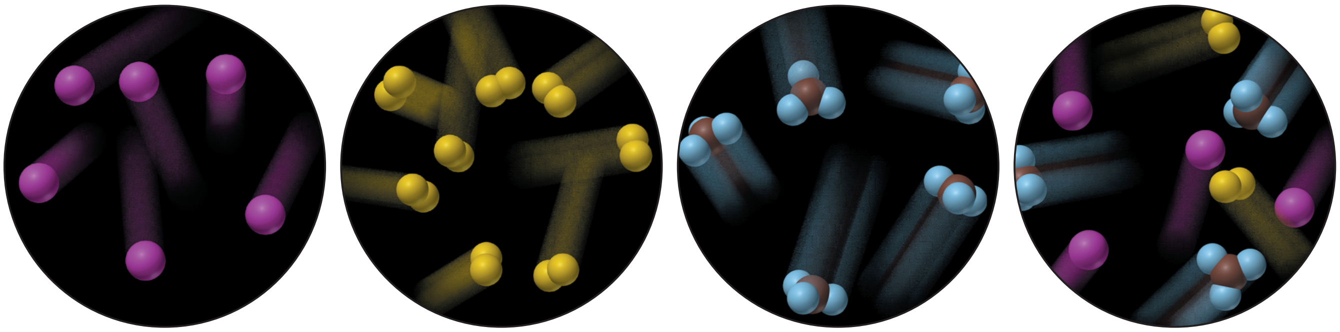


Figure 2.1 Elements, compounds, and mixtures on the atomic scale. **A,** Most elements consist of a large collection of identical atoms. **B,** Some elements occur as molecules. **C,** A molecule of a compound consists of characteristic numbers of atoms of two or more elements chemically bound together. **D,** A mixture contains the individual units of two or more elements and/or compounds that are physically intermingled. The samples shown here are gases, but elements, compounds, and mixtures occur as liquids and solids also.

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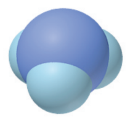
Most elements exist in nature as populations of atoms. Figure 2.1A shows atoms of a gaseous element such as neon. However, several elements occur naturally as molecules: a **molecule** is an independent structure consisting of two or more atoms chemically bound together (Figure 2.1B). Elemental oxygen, for example, occurs in air as *diatomic* (two-atom) molecules.

A **compound** is a type of matter composed of *two or more different elements that are chemically bound together.* Be sure you understand that the elements in a compound are not just mixed together; rather, their atoms have joined chemically (Figure 2.1C). Ammonia, water, and carbon dioxide are some common compounds. One defining feature of a compound is that *the elements are present in fixed parts by mass* (fixed mass ratio). Because of this fixed composition, *a compound is also considered a substance*. Any sample of the compound has the same fixed parts by mass because each of its molecules consists of *fixed numbers* of atoms of the component elements. For example, any sample of ammonia is 14 parts nitrogen by mass plus 3 parts hydrogen by mass. Since 1 nitrogen atom has 14 times the mass of 1 hydrogen atom, a molecule of ammonia must consist of 1 nitrogen atom for every 3 hydrogen atoms:

Ammonia is 14 parts N and 3 parts H by mass.
1 N atom has 14 times the mass of 1 H atom.
Therefore, ammonia has 1 N atom for every 3 H atoms.



Another defining feature of a compound is that *its properties are different from those of its component elements*. Table 2.1 shows a striking example. Soft, silvery sodium metal and yellow-green, poisonous chlorine gas have very different properties from the compound they form—white, crystalline sodium chloride, or common table salt! Unlike an element, a compound *can* be broken down into simpler substances—its component elements. For example, an electric current breaks down molten sodium chloride into metallic sodium and chlorine gas. Note that this breakdown is a *chemical change,* not a physical one.

Figure 2.1D depicts a **mixture,** a group of two or more substances (elements and/or compounds) that are physically intermingled. In contrast to a compound,

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Table 2.1 Some Properties of Sodium, Chlorine, and Sodium Chloride

Property	Sodium	+	Chlorine	→	Sodium Chloride	
Melting point	97.8°C		−101°C		801°C	
Boiling point	881.4°C		−34°C		1413°C	
Color	Silvery		Yellow-green		Colorless (white)	
Density	0.97 g/cm^3		0.0032 g/cm^3		2.16 g/cm^3	
Behavior in water	Reacts		Dissolves slightly		Dissolves freely	

the components of a mixture ***can*** *vary in their parts by mass.* Because its composition is not fixed, a mixture is *not* a substance. A mixture of the two compounds sodium chloride and water, for example, can have many different parts by mass of salt to water. Because the components are physically mixed, not chemically combined, a mixture at the atomic scale is merely a group of the individual units that make up its component elements and/or compounds. Therefore, *a mixture retains many of the properties of its components.* Saltwater, for instance, is colorless like water and tastes salty like sodium chloride. Unlike compounds, mixtures can be separated into their components by *physical changes;* chemical changes are not needed. For example, the water in saltwater can be boiled off, a physical process that leaves behind the solid sodium chloride. The following sample problem will help differentiate these types of matter.

SAMPLE PROBLEM 2.1 Distinguishing Elements, Compounds, and Mixtures at the Atomic Scale

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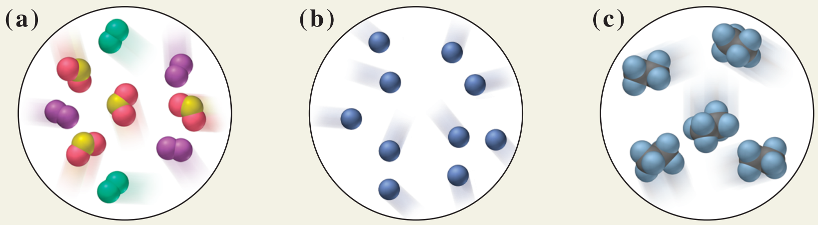
PROBLEM The scenes below represent an atomic-scale view of three samples of matter:

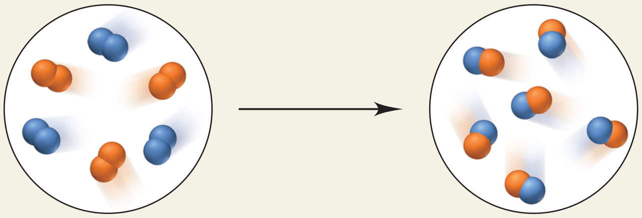


Describe each sample as an element, compound, or mixture.

PLAN From depictions of the samples, we have to determine the type of matter by examining the component particles. If a sample contains only one type of particle, it is either an element or a compound; if it contains more than one type, it is a mixture. Particles of an element have only one kind of atom (one color), and particles of a compound have two or more kinds of atoms.

SOLUTION (**a**) This sample is a mixture: there are three different types of particles, two types contain only one kind of atom, either green or purple, so they are elements, and the third type contains two red atoms for every one yellow, so it is a compound. (**b**) This sample is an element: it consists of only blue atoms, (**c**) This sample is a compound: it consists of molecules that each have two black and six blue atoms.

FOLLOW-UP PROBLEM 2.1 Describe the following reaction in terms of elements, compounds, and mixtures.



Section Summary

All matter exists as either elements, compounds, or mixtures. • Elements and compounds are referred to as *substances* because their compositions are fixed. An element consists of only one type of atom. A compound contains two or more elements in chemical combination and exhibits different properties from its component elements. The elements of a compound occur in fixed parts by mass because each unit of the compound has fixed numbers of each type of atom. • A mixture consists of two or more substances mixed together, not chemically combined. The components retain their individual properties and can be present in any proportion.

Literature Clark, Roy W. "Why Do Chemists Weigh Things in Grams Instead of Gamus?" *Chem. Educator* 7 **2002**, *4*, 192–194.

2.2 THE OBSERVATIONS THAT LED TO AN ATOMIC VIEW OF MATTER

Any model of the composition of matter had to explain two extremely important chemical observations that were well established by the end of the 18th century: the *law of mass conservation* and the *law of definite (or constant) composition*. As you'll see, John Dalton's atomic theory explained these laws and another observation now known as the *law of multiple proportions*.

Mass Conservation

The most fundamental chemical observation of the 18th century was the **law of mass conservation:** *the total mass of substances does not change during a chemical reaction.* The *number* of substances may change and, by definition, their properties must, but the *total amount* of matter remains constant. Lavoisier had first stated this law on the basis of his combustion experiments. Figure 2.2 illustrates mass conservation in a reaction that occurs in water.

Even in a complex biochemical change, such as the metabolism of the sugar glucose, which involves many reactions, mass is conserved:

$$\text{180 g glucose} + \text{192 g oxygen gas} \longrightarrow \text{264 g carbon dioxide} + \text{108 g water}$$
$$\text{372 g material before change} \longrightarrow \text{372 g material after change}$$

Mass conservation means that, based on all chemical experience, *matter cannot be created or destroyed.* (As you'll see later, however, mass *does* change in nuclear reactions.)

Immeasurable Changes in Mass
Based on the work of Albert Einstein (1879–1955), we now know that some mass *does* change into energy during a chemical reaction. But the amount is too small to measure, even by the best modern balance. For example, when 100 g of carbon burns in oxygen, carbon dioxide is formed, and only 0.000000036 g (3.6×10^{-8} g) of mass is converted to energy. The energy yields of chemical reactions are relatively so small that, for all practical purposes, mass *is* conserved. As you'll see later, however, energy changes in nuclear reactions are so large that mass changes are measured easily.

Definite Composition

Another fundamental chemical observation is summarized as the **law of definite (or constant) composition:** *no matter what its source, a particular compound is composed of the same elements in the same parts (fractions) by mass.* The **fraction by mass (mass fraction)** is that part of the compound's mass that each element contributes. It is obtained by dividing the mass of each element by the total mass of compound. The **percent by mass (mass percent, mass %)** is the fraction by mass expressed as a percentage.

Let's see what these ideas mean in terms of a box of marbles *(right)*. The box contains three types of marbles: yellow marbles weigh 1.0 g each, purple marbles 2.0 g each, and red marbles 3.0 g each. Each type makes up a fraction of the total mass of marbles, 16.0 g. The *mass fraction*

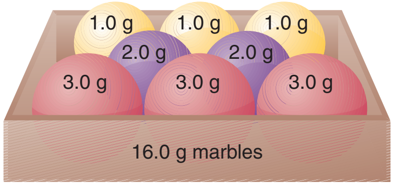


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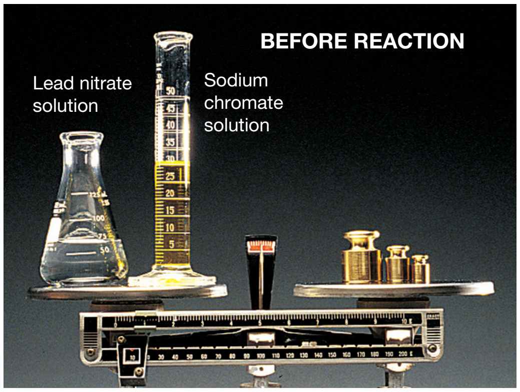
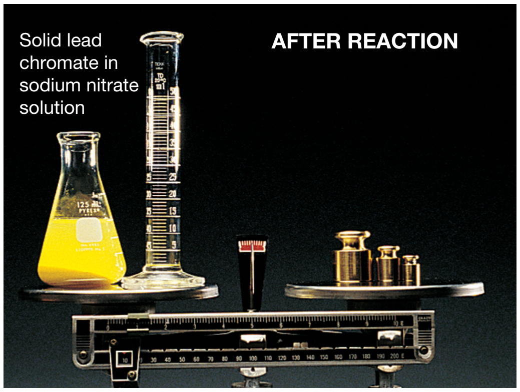
Figure 2.2 The law of mass conservation: mass remains constant during a chemical reaction. The total mass of lead nitrate solution and sodium chromate solution before they react **(A)** is the same as the total mass after they have reacted **(B)** to form lead chromate (yellow solid) and sodium nitrate solution.

A



B



of the yellow marbles is their number times their mass divided by the total mass: $(3 \times 1.0\text{ g})/16.0\text{ g} = 0.19$. The *mass percent* (parts per 100 parts) of the yellow marbles is $0.19 \times 100 = 19\%$ by mass. The purple marbles have a mass fraction of 0.25 and are 25% of the total by mass, and the red marbles have a mass fraction of 0.56 and are 56% by mass. Similarly, in a compound, each element has a *fixed* mass fraction (and mass percent).

Consider calcium carbonate, the major compound in marble. It is composed of three elements—calcium, carbon, and oxygen—and each is present in a fixed fraction (or percent) by mass. The following results are obtained for the elemental mass composition of 20.0 g of calcium carbonate (for example, 8.0 g of calcium/20.0 g = 0.40 parts of calcium):

Analysis by Mass (grams/20.0 g)	Mass Fraction (parts/1.00 part)	Percent by Mass (parts/100 parts)
8.0 g calcium	0.40 calcium	40% calcium
2.4 g carbon	0.12 carbon	12% carbon
9.6 g oxygen	0.48 oxygen	48% oxygen
20.0 g	1.00 part by mass	100% by mass

As you can see, the sum of the mass fractions (or mass percents) equals 1.00 part (or 100%) by mass. The law of definite composition tells us that pure samples of calcium carbonate, no matter where they come from, always contain these elements in the same percents by mass (Figure 2.3).

Because a given element always constitutes the same mass fraction of a given compound, we can use that mass fraction to find the actual mass of the element in any sample of the compound:

$$\text{Mass of element} = \text{mass of compound} \times \frac{\text{part by mass of element}}{\text{one part by mass of compound}}$$

Or, more simply, because mass analysis tells us the parts by mass, we can use that ratio directly with *any* mass unit and skip the need to find the mass fraction first:

$$\text{Mass of element in sample} = \text{mass of compound in sample} \times \frac{\text{mass of element in compound}}{\text{mass of compound}} \qquad \textbf{(2.1)}$$

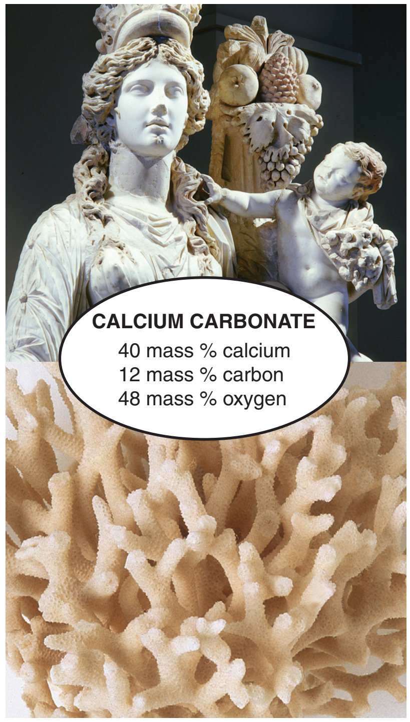


Figure 2.3 The law of definite composition. Calcium carbonate is found naturally in many forms, including marble *(top)*, coral *(bottom)*, chalk, and seashells. The mass percents of its component elements do not change regardless of the compound's source.

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SAMPLE PROBLEM 2.2 Calculating the Mass of an Element in a Compound

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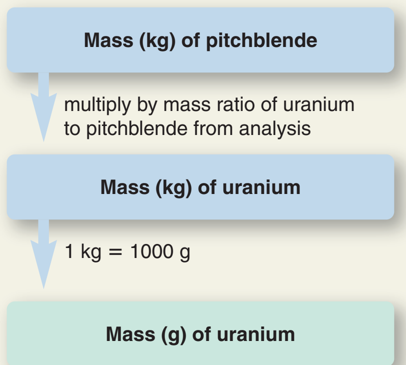
PROBLEM Pitchblende is the most commercially important compound of uranium. Analysis shows that 84.2 g of pitchblende contains 71.4 g of uranium, with oxygen as the only other element. How many grams of uranium can be obtained from 102 kg of pitchblende?

PLAN We have to find the mass of uranium in a known mass of pitchblende, given the mass of uranium in a different mass of pitchblende. The mass ratio of uranium/pitchblende is the same for any sample of pitchblende. Therefore, as shown by Equation 2.1, we multiply the mass (in kg) of pitchblende by the ratio of uranium to pitchblende that we construct from the mass analysis. This gives the mass (in kg) of uranium, and we just convert kilograms to grams.



SOLUTION Finding the mass (kg) of uranium in 102 kg of pitchblende:

$$\text{Mass (kg) of uranium} = \text{mass (kg) of pitchblende} \times \frac{\text{mass (kg) of uranium in pitchblende}}{\text{mass (kg) of pitchblende}}$$

$$\text{Mass (kg) of uranium} = 102\ \cancel{\text{kg pitchblende}} \times \frac{71.4\text{ kg uranium}}{84.2\ \cancel{\text{kg pitchblende}}} = 86.5\text{ kg uranium}$$

Converting the mass of uranium from kg to g:

$$\text{Mass (g) of uranium} = 86.5\ \cancel{\text{kg}}\text{ uranium} \times \frac{1000\text{ g}}{1\ \cancel{\text{kg}}} = \boxed{8.65 \times 10^4\text{ g uranium}}$$

CHECK The analysis showed that most of the mass of pitchblende is due to uranium, so the large mass of uranium makes sense. Rounding off to check the math gives:

$$\sim 100 \text{ kg pitchblende} \times \frac{70}{85} = 82 \text{ kg uranium}$$

FOLLOW-UP PROBLEM 2.2 How many metric tons (t) of oxygen are combined in a sample of pitchblende that contains 2.3 t of uranium? (*Hint:* Remember that oxygen is the only other element present.)

Demo Shakhashiri, Bassam Z. "Reaction of Iron and Sulfur," *Chemical Demonstrations, A Handbook for Teachers of Chemistry,* Vol. 1 (The University of Wisconsin Press, Madison, 1985): 55.

Multiple Proportions

Dalton described a phenomenon that occurs when two elements form more than one compound. His observation is now called the **law of multiple proportions:** *if elements A and B react to form two compounds, the different masses of B that combine with a fixed mass of A can be expressed as a ratio of small whole numbers*. Consider two compounds that form from carbon and oxygen; for now, let's call them carbon oxides I and II. They have very different properties. For example, measured at the same temperature and pressure, the density of carbon oxide I is 1.25 g/L, whereas that of II is 1.98 g/L. Moreover, I is poisonous and flammable, but II is not. Analysis shows that their compositions by mass are

Carbon oxide I: 57.1 mass % oxygen and 42.9 mass % carbon
Carbon oxide II: 72.7 mass % oxygen and 27.3 mass % carbon

To see the phenomenon of multiple proportions, we use the mass percents of oxygen and of carbon in each compound to find the masses of these elements in a given mass, for example, 100 g, of each compound. Then we divide the mass of oxygen by the mass of carbon in each compound to obtain the mass of oxygen that combines with a fixed mass of carbon:

	Carbon Oxide I	Carbon Oxide II
g oxygen/100 g compound	57.1	72.7
g carbon/100 g compound	42.9	27.3
g oxygen/g carbon	$\frac{57.1}{42.9} = 1.33$	$\frac{72.7}{27.3} = 2.66$

If we then divide the grams of oxygen per gram of carbon in II by that in I, we obtain a ratio of small whole numbers:

$$\frac{2.66 \text{ g oxygen/g carbon in II}}{1.33 \text{ g oxygen/g carbon in I}} = \frac{2}{1}$$

The law of multiple proportions tells us that in two compounds of the same elements, the mass fraction of one element relative to the other element changes in *increments based on ratios of small whole numbers*. In this case, the ratio is 2:1—for a given mass of carbon, II contains *2 times* as much oxygen as I, not 1.583 times, 1.716 times, or any other intermediate amount. As you'll see next, Dalton's theory allows us to explain the composition of carbon oxides I and II on the atomic scale.

Section Summary

Three fundamental observations are known as the mass laws. The law of mass conservation states that the total mass remains constant during a chemical reaction. • The law of definite composition states that any sample of a given compound has the same elements present in the same parts by mass. • The law of multiple proportions states that, in different compounds of the same elements, the masses of one element that combine with a fixed mass of the other can be expressed as a ratio of small whole numbers.

2.3 DALTON'S ATOMIC THEORY

With 200 years of hindsight, it may be easy to see how the mass laws could be explained by an atomic model—matter existing in indestructible units, each with a particular mass—but it was a major breakthrough in 1808 when John Dalton (1766–1844) presented his atomic theory of matter in *A New System of Chemical Philosophy.*

Postulates of the Atomic Theory

Dalton expressed his theory in a series of postulates. Like most great thinkers, Dalton incorporated the ideas of others into his own to create the new theory. As we go through the postulates, which are presented here in modern terms, let's see which were original and which came from others. (Later, we can examine the key differences between Dalton's postulates and our present understanding.)

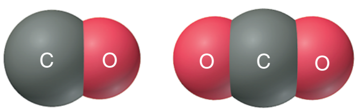
1. All matter consists of **atoms,** tiny indivisible particles of an element that cannot be created or destroyed. (Derives from the "eternal, indestructible atoms" proposed by Democritus more than 2000 years earlier and conforms to mass conservation as stated by Lavoisier.)
2. Atoms of one element *cannot* be converted into atoms of another element. In chemical reactions, the atoms of the original substances recombine to form different substances. (Rejects the alchemical belief in the magical transmutation of elements.)
3. Atoms of an element are identical in mass and other properties and are different from atoms of any other element. (Contains Dalton's major new ideas: unique mass and properties for all the atoms of a given element.)
4. Compounds result from the chemical combination of a specific ratio of atoms of different elements. (Follows directly from the law of definite composition.)

How the Theory Explains the Mass Laws

Let's see how Dalton's postulates explain the mass laws:

- *Mass conservation.* Atoms cannot be created or destroyed (postulate 1) or converted into other types of atoms (postulate 2). Since each type of atom has a fixed mass (postulate 3), a chemical reaction, in which atoms are just combined differently with each other, cannot possibly result in a mass change.
- *Definite composition.* A compound is a combination of a *specific* ratio of different atoms (postulate 4), each of which has a particular mass (postulate 3). Thus, each element in a compound constitutes a fixed fraction of the total mass.
- *Multiple proportions.* Atoms of an element have the same mass (postulate 3) and are indivisible (postulate 1). The masses of element B that combine with a fixed mass of element A give a small, whole-number ratio because different numbers of B atoms combine with each A atom in different compounds.

The *simplest* arrangement consistent with the mass data for carbon oxides I and II in our earlier example is that one atom of oxygen combines with one atom of carbon in compound I (carbon monoxide) and that two atoms of oxygen combine with one atom of carbon in compound II (carbon dioxide):



Carbon oxide I (carbon monoxide)   Carbon oxide II (carbon dioxide)

Let's work through a sample problem that reviews the mass laws.

Dalton's Revival of Atomism Although John Dalton, the son of a poor weaver, had no formal education, he established one of the most powerful concepts in science. Dalton began teaching science at 12 years of age, and later studied color blindness, a personal affliction still known as *daltonism.* In 1787, he began his life's work in meteorology, recording daily weather data until his death 57 years later. His studies on humidity and dew point led to a key discovery about the behavior of gases (Section 5.4) and eventually to his atomic theory. In 1803, he stated, "I am nearly persuaded that [the mixing of gases and their solubility in water] depends upon the mass and number of the ultimate particles. . . . An enquiry into the relative masses of [these] particles of bodies is a subject . . . I have lately been prosecuting . . . with remarkable success." The atomic theory was published 5 years later.

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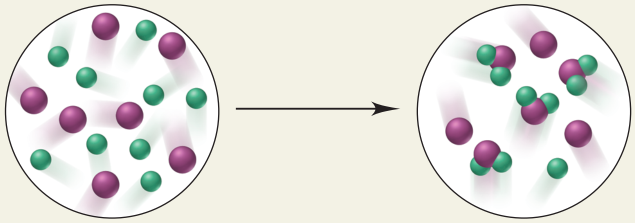
SAMPLE PROBLEM 2.3 Visualizing the Mass Laws

PROBLEM The scenes below represent an atomic-scale view of a chemical reaction:



Which of the mass laws—mass conservation, definite composition, or multiple proportions—is (are) illustrated?

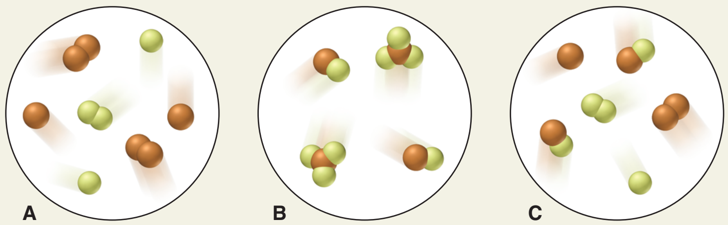
PLAN From the depictions, we note the number, color, and combinations of atoms (spheres) to see which mass laws pertain. If the numbers of each atom are the same before and after the reaction, the total mass did not change (mass conservation). If a compound forms that always has the same atom ratio, the elements are present in fixed parts by mass (definite composition). When the same elements form different compounds and the ratio of the atoms of one element that combine with one atom of the other element is a small whole number, the ratio of their masses is a small whole number as well (multiple proportions).

SOLUTION There are seven purple and nine green atoms in each circle, so mass is conserved.

The compound formed has one purple and two green atoms, so it has definite composition.

Only one compound forms, so the law of multiple proportions does not pertain.

FOLLOW-UP PROBLEM 2.3 Which sample(s) best display(s) the fact that compounds of bromine (orange) and fluorine (yellow) exhibit the law of multiple proportions? Explain.



A B C

Section Summary

Dalton's atomic theory explained the mass laws by proposing that all matter consists of indivisible, unchangeable atoms of fixed, unique mass. • Mass is conserved during a reaction because atoms form new combinations. • Each compound has a fixed mass fraction of each of its elements because it is composed of a fixed number of each type of atom. • Different compounds of the same elements exhibit multiple proportions because they each consist of whole atoms.

2.4 THE OBSERVATIONS THAT LED TO THE NUCLEAR ATOM MODEL

After publication of the atomic theory, investigators tried to determine the masses of atoms from the mass fractions of elements in compounds. Because an individual atom is so small, the mass of the atoms of one element was determined *relative* to the mass of the atoms of another element, based on a mass standard. Dalton's model was crucial because it originated the idea that masses of reacting elements could be explained in terms of atoms. However, the model did not explain why atoms bond as they do: for example, why do two, and not three, hydrogen atoms bond with one oxygen atom in water? Also, Dalton's "billiard ball" atom did not account for the charged particles observed in later experiments. Clearly, a more complex atomic model was needed. ●

Basic research into the nature of electricity eventually led to the discovery of *electrons,* negatively charged particles that are part of all atoms. Soon thereafter,

● **Atoms? Humbug!** Rarely does a major new concept receive unanimous acceptance. Despite the atomic theory's impact, several major scientists denied the existence of atoms for another century. In 1877, Adolf Kolbe, an eminent organic chemist, said, "[Dalton's atoms are] . . . no more than stupid hallucinations . . . mere table-tapping and supernatural explanations." The influential physicist Ernst Mach believed that scientists should look at facts, not hypothetical entities such as atoms. It was not until 1908 that the famous chemist and outspoken opponent of atomism Wilhelm Ostwald wrote, "I am now convinced [by recent] experimental evidence of the discrete or grained nature of matter, which the atomic hypothesis sought in vain for hundreds and thousands of years." He was referring to the discovery of the electron.

other experiments revealed that the atom has a *nucleus*—a tiny, central core of mass and positive charge. In this section, we examine some key experiments that led to our current model of the atom.

Discovery of the Electron and Its Properties

Nineteenth-century investigators of electricity knew that matter and electric charge were somehow related. When amber is rubbed with fur, or glass with silk, positive and negative charges form—the same charges that make your hair crackle and cling to your comb on a dry day. They also knew that an electric current could decompose certain compounds into their elements.

What they did not know, however, was what an electric current itself might consist of. Some investigators tried passing current from a high-voltage source through nearly evacuated glass tubes fitted with metal electrodes that were sealed in place and connected to an external source of electricity. When the power was turned on, a "ray" could be seen striking the phosphor-coated end of the tube and emitting a glowing spot of light. The rays were called **cathode rays** because they originated at the negative electrode (cathode) and moved to the positive electrode (anode). Cathode rays typically travel in a straight line, but in a magnetic field the path is bent, indicating that the particles are charged, and in an electric field the path bends toward the positive plate. The ray is identical no matter what metal is used as the cathode (Figure 2.4). It was concluded that cathode rays consist of negatively charged particles found in all matter. The rays appear when these particles collide with the few remaining gas molecules in the evacuated tube. ● Cathode ray particles were later named *electrons*.

● **The Familiar Glow of Colliding Particles** The electric and magnetic properties of charged particles that collide with gas particles or hit a phosphor-coated screen have familiar applications. A "neon" sign glows because electrons collide with the gas particles in the tube, causing them to give off light. An aurora display occurs when Earth's magnetic field bends streams of charged particles coming from the Sun, which then collide with gases in the atmosphere. In a television tube or computer monitor, the cathode ray passes back and forth over the coated screen, creating a pattern that the eye sees as a picture.

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Animation: Cathode Ray Tube
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Animations Center, Chapter 2

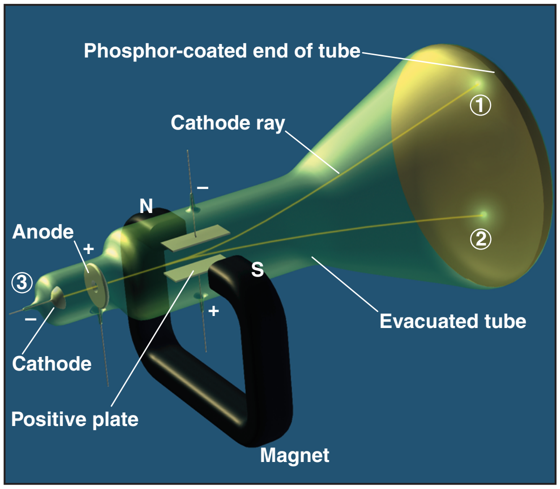

Figure 2.4 Experiments to determine the properties of cathode rays. A cathode ray forms when high voltage is applied to a partially evacuated tube. The ray passes through a hole in the anode and hits the coated end of the tube to produce a glow.

OBSERVATION	CONCLUSION
1. Ray bends in magnetic field	Consists of charged particles
2. Ray bends toward positive plate in electric field	Consists of negative particles
3. Ray is identical for any cathode	Particles found in all matter

In 1897, the British physicist J. J. Thomson (1856–1940) used magnetic and electric fields to measure the ratio of the cathode ray particle's mass to its charge. By comparing this value with the mass/charge ratio for the lightest charged particle in solution, Thomson estimated that the cathode ray particle weighed less than $\frac{1}{1000}$ as much as hydrogen, the lightest atom! He was shocked because this implied that, contrary to Dalton's atomic theory, *atoms are divisible into even smaller particles*. Thomson concluded, "We have in the cathode rays matter in a new state, . . . in which the subdivision of matter is carried much further . . . ; this matter being the substance from which the chemical elements are built up." Fellow scientists reacted with disbelief, and some even thought he was joking.

In 1909, the American physicist Robert Millikan (1868–1953) measured the *charge* of the electron. He did so by observing the movement of tiny droplets of the "highest grade clock oil" in an apparatus that contained electrically charged

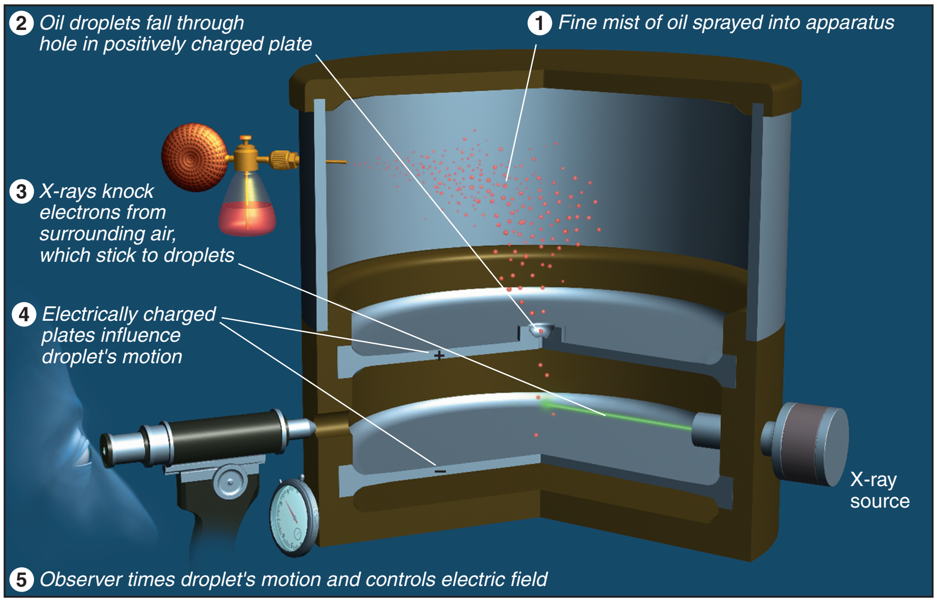

Figure 2.5 Millikan's oil-drop experiment for measuring an electron's charge. The motion of a given oil droplet depends on the variation in electric field and the total charge on the droplet, which depends in turn on the number of attached electrons. Millikan reasoned that the total charge must be some whole-number multiple of the charge of the electron.

Transparency/ARIS Presentation Center, Active Art Library

Literature Pearson, Earl F. "Millikan: Good to the Last (Oil) Drop," *J. Chem. Educ.* **2006,** *83,* 1312A.

Literature Eckey, Doris. "A Millikan Oil Drop Analogy," *J. Chem. Educ.* **1996,** *73,* 237.

Animation: Millikan Oil Drop
ARIS Presentation Center
Chapter 2

plates and an x-ray source (Figure 2.5). Here is a description of the basis of the experiment: X-rays knocked electrons from gas molecules in the air, and as an oil droplet fell through a hole in the positive (upper) plate, the electrons stuck to the drop, giving it a negative charge. With the electric field off, Millikan measured the mass of the droplet from its rate of fall. By turning on the field and varying its strength, he could make the drop fall more slowly, rise, or pause suspended. From these data, Millikan calculated the total charge of the droplet.

After studying many droplets, Millikan calculated that the various charges of the droplets were always some *whole-number multiple of a minimum charge.* He reasoned that different oil droplets picked up different numbers of electrons, so this minimum charge must be that of the electron itself. The value that he calculated a century ago was within 1% of the modern value of the electron's charge, -1.602×10^{-19} C (C stands for *coulomb,* the SI unit of charge). Using the electron's mass/charge ratio from work by Thomson and others and this value for the electron's charge, let's calculate the electron's *extremely* small mass the way Millikan did:

$$\text{Mass of electron} = \frac{\text{mass}}{\text{charge}} \times \text{charge} = \left(-5.686\times10^{-12}\,\frac{\text{kg}}{\cancel{\text{C}}}\right)(-1.602\times10^{-19}\,\cancel{\text{C}})$$
$$= 9.109\times10^{-31}\text{ kg} = 9.109\times10^{-28}\text{ g}$$

Discovery of the Atomic Nucleus

Clearly, the properties of the electron posed problems about the inner structure of atoms. If everyday matter is electrically neutral, the atoms that make it up must be neutral also. But if atoms contain negatively charged electrons, what positive charges balance them? And if an electron has such an incredibly tiny mass, what accounts for an atom's much larger mass? To address these issues, Thomson proposed a model of a spherical atom composed of diffuse, positively charged matter, in which electrons were embedded like "raisins in a plum pudding."

Near the turn of the 20th century, French scientists discovered radioactivity, the emission of particles and/or radiation from atoms of certain elements. Just a few years later, in 1910, the New Zealand–born physicist Ernest Rutherford (1871–1937) used one type of radioactive particle in a series of experiments that solved this dilemma of atomic structure.

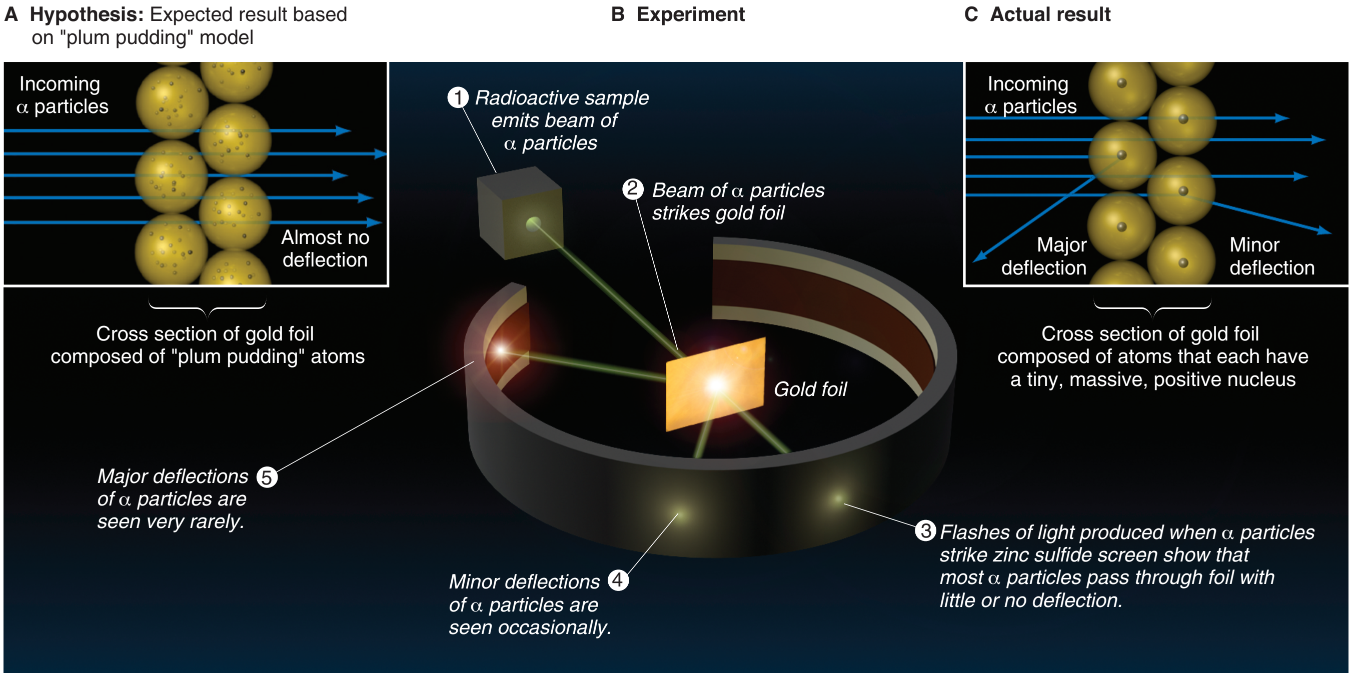

Figure 2.6 Rutherford's α-scattering experiment and discovery of the atomic nucleus. **A,** HYPOTHESIS: Atoms consist of electrons embedded in diffuse, positively charged matter, so the speeding α particles should pass through the gold foil with, at most, minor deflections. **B,** EXPERIMENT: α Particles emit a flash of light when they pass through the gold atoms and hit a phosphor-coated screen. **C,** RESULTS: Occasional minor deflections and very infrequent major deflections are seen. This means very high mass and positive charge are concentrated in a small region within the atom, the nucleus.

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Animation: Rutherford's Experiment
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Chapter 2

Figure 2.6 is a three-part representation of Rutherford's experiment. Tiny, dense, positively charged alpha (α) particles emitted from radium were aimed, like minute projectiles, at thin gold foil. The figure illustrates (A) the "plum pudding" hypothesis, (B) the apparatus used to measure the deflection (scattering) of the α particles from the light flashes created when the particles struck a circular, coated screen, and (C) the actual result.

With Thomson's model in mind, Rutherford expected only minor, if any, deflections of the α particles because they should act as tiny, dense, positively charged "bullets" and go right through the gold atoms. According to the model, one of the embedded electrons could not deflect an α particle any more than a Ping-Pong ball could deflect a speeding baseball. Initial results confirmed this, but soon the unexpected happened. As Rutherford recalled: "Then I remember two or three days later Geiger [one of his coworkers] coming to me in great excitement and saying, 'We have been able to get some of the α particles coming backwards . . .' It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you."

The data showed that very few α particles were deflected at all, and that only 1 in 20,000 was deflected by more than 90° ("coming backwards"). It seemed that these few α particles were being repelled by something small, dense, and positive within the gold atoms. From the mass, charge, and velocity of the α particles, the frequency of these large-angle deflections, and the properties of electrons, Rutherford calculated that *an atom is mostly space occupied by electrons,* but in the center of that space is a tiny region, which he called the **nucleus,** that contains *all the positive charge and essentially all the mass of the atom.* He proposed that positive particles lay within the nucleus and called them *protons,* and then he calculated the magnitude of the nuclear charge with remarkable accuracy. Rutherford's model explained the charged nature of matter, but it could not account for all the atom's mass. After more than 20 years, this issue was resolved when, in 1932, James Chadwick discovered the *neutron,* an uncharged dense particle that also resides in the nucleus.

Section Summary

Several major discoveries at the turn of the 20th century led to our current model of atomic structure. • Cathode rays were shown to consist of negative particles (electrons) that exist in all matter. J. J. Thomson measured their mass/charge ratio and concluded that they are much smaller and lighter than atoms. • Robert Millikan determined the charge of the electron, which he combined with other data to calculate its mass. • Ernest Rutherford proposed that atoms consist of a tiny, massive, positive nucleus surrounded by electrons.

2.5 THE ATOMIC THEORY TODAY

For over 200 years, scientists have known that all matter consists of atoms, and they have learned astonishing things about them. Dalton's tiny indivisible particles have given way to atoms with "fuzzy," indistinct boundaries and an elaborate internal architecture of subatomic particles. In this section, we examine our current model and begin to see how the properties of subatomic particles affect the properties of atoms. Then we'll see how Dalton's theory stands up today.

Structure of the Atom

An *atom* is an electrically neutral, spherical entity composed of a positively charged central nucleus surrounded by one or more negatively charged electrons (Figure 2.7). The electrons move rapidly within the available atomic volume, held there by the attraction of the nucleus. The nucleus is incredibly dense: it contributes 99.97% of the atom's mass but occupies only about 1 quadrillionth of its volume. (A nucleus the size of a period on this page would weigh about 100 tons, as much as 50 cars!) An atom's diameter ($\sim10^{-10}$ m) is about 100,000 times the diameter of its nucleus ($\sim10^{-15}$ m).

Literature Peake, Barrie M. "The Discovery of the Electron, Proton, and Neutron," *J. Chem. Educ.* **1989,** *66,* 738.

An atomic nucleus consists of protons and neutrons (the only exception is the simplest hydrogen nucleus, which is a single proton). The **proton (p^+)** has a positive charge, and the **neutron (n^0)** has no charge; thus, the positive charge of the nucleus results from its protons. The *magnitude* of charge possessed by a proton is equal to that of an **electron (e^-),** but the *signs* of the charges are opposite. *An atom is neutral because the number of protons in the nucleus equals the number of electrons surrounding the nucleus.* Some properties of these three subatomic particles are listed in Table 2.2.

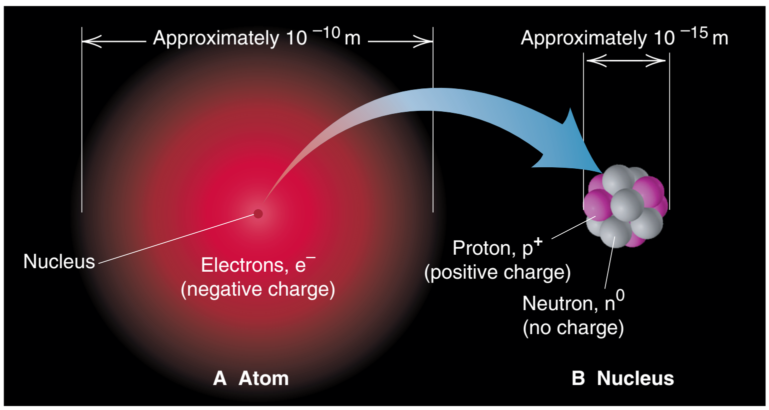

Figure 2.7 General features of the atom. A, A "cloud" of rapidly moving, negatively charged electrons occupies virtually all the atomic volume and surrounds the tiny, central nucleus. **B,** The nucleus contains virtually all the mass of the atom and consists of positively charged protons and uncharged neutrons. If the nucleus were actually the size in the figure (~1 cm across), the atom would be about 1000 m (1 km) across.

Animation: Alpha, Beta, and Gamma Rays
ARIS Presentation Center
Chapter 2

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Table 2.2 Properties of the Three Key Subatomic Particles

Name (Symbol)	Charge: Relative	Charge: Absolute (C)*	Mass: Relative (amu)†	Mass: Absolute (g)	Location in Atom
Proton (p^+)	1+	$+1.60218\times10^{-19}$	1.00727	1.67262×10^{-24}	Nucleus
Neutron (n^0)	0	0	1.00866	1.67493×10^{-24}	Nucleus
Electron (e^-)	1−	-1.60218×10^{-19}	0.00054858	9.10939×10^{-28}	Outside nucleus

*The coulomb (C) is the SI unit of charge.
†The atomic mass unit (amu) equals 1.66054×10^{-24} g; discussed later in this section.

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Atomic Number, Mass Number, and Atomic Symbol

The **atomic number (*Z*)** of an element equals the number of protons in the nucleus of each of its atoms. *All atoms of a particular element have the same atomic number, and each element has a different atomic number from that of any other element.* All carbon atoms ($Z = 6$) have 6 protons, all oxygen atoms ($Z = 8$) have 8 protons, and all uranium atoms ($Z = 92$) have 92 protons. There are currently 116 known elements, of which 90 occur in nature; the remaining 26 have been synthesized by nuclear scientists.

The total number of protons and neutrons in the nucleus of an atom is its **mass number (*A*).** Each proton and each neutron contributes one unit to the mass number. Thus, a carbon atom with 6 protons and 6 neutrons in its nucleus has a mass number of 12, and a uranium atom with 92 protons and 146 neutrons in its nucleus has a mass number of 238.

The nuclear mass number and charge are often written with the **atomic symbol** (or *element symbol*). Every element has a symbol based on its English, Latin, or Greek name, such as C for carbon, O for oxygen, S for sulfur, and Na for sodium (Latin *natrium*). The atomic number (*Z*) is written as a left *sub*script and the mass number (*A*) as a left *super*script to the symbol, so element X would be $^{A}_{Z}\text{X}$. Since the mass number is the sum of protons and neutrons, the number of neutrons (*N*) equals the mass number minus the atomic number:

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}, \quad \text{or} \quad N = A - Z \qquad \textbf{(2.2)}$$

Thus, a chlorine atom, which is symbolized as $^{35}_{17}\text{Cl}$, has $A = 35$, $Z = 17$, and $N = 35 - 17 = 18$. Each element has its own atomic number, so we know the atomic number from the symbol. For example, every carbon atom has 6 protons. Therefore, instead of writing $^{12}_{6}\text{C}$ for carbon with mass number 12, we can write ^{12}C (spoken "carbon twelve"), with $Z = 6$ understood. Another way to write this atom is carbon-12.

Isotopes and Atomic Masses of the Elements

All atoms of an element are identical in atomic number but not in mass number. **Isotopes** of an element are atoms that have *different numbers of neutrons* and therefore different mass numbers. For example, all carbon atoms ($Z = 6$) have 6 protons and 6 electrons, but only 98.89% of naturally occurring carbon atoms have 6 neutrons in the nucleus ($A = 12$). A small percentage (1.11%) have 7 neutrons in the nucleus ($A = 13$), and even fewer (less than 0.01%) have 8 ($A = 14$). These are carbon's three naturally occurring isotopes—^{12}C, ^{13}C, and ^{14}C. Five other carbon isotopes—^{9}C, ^{10}C, ^{11}C, ^{15}C, and ^{16}C—have been created in the laboratory. Figure 2.8 depicts the atomic number, mass number, and symbol for four atoms, two of which are isotopes of the same element.

A key point is that the chemical properties of an element are primarily determined by the number of electrons, so *all isotopes of an element have nearly identical chemical behavior,* even though they have different masses.

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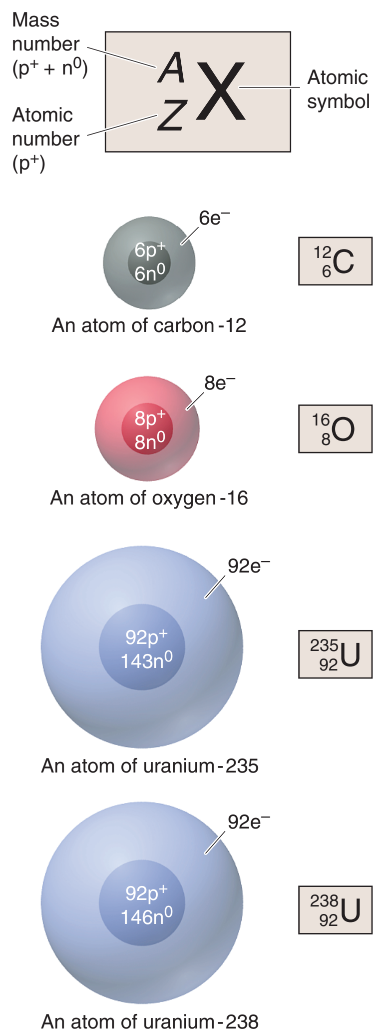

Figure 2.8 Depicting the atom. Atoms of carbon-12, oxygen-16, uranium-235, and uranium-238 are shown (nuclei not drawn to scale) with their symbolic representations. The sum of the number of protons (*Z*) and the number of neutrons (*N*) equals the mass number (*A*). An atom is neutral, so the number of protons in the nucleus equals the number of electrons around the nucleus. The two uranium atoms are isotopes of the element.

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SAMPLE PROBLEM 2.4 Determining the Number of Subatomic Particles in the Isotopes of an Element

PROBLEM Silicon (Si) is essential to the computer industry as a major component of semiconductor chips. It has three naturally occurring isotopes: ^{28}Si, ^{29}Si, and ^{30}Si. Determine the numbers of protons, neutrons, and electrons in each silicon isotope.

PLAN The mass number (A) of each of the three isotopes is given, so we know the sum of protons and neutrons. From the elements list on the text's inside front cover, we find the atomic number (Z, number of protons), which equals the number of electrons. We obtain the number of neutrons from Equation 2.2.

SOLUTION From the elements list, the atomic number of silicon is 14. Therefore,

$$^{28}\text{Si has } 14p^+,\ 14e^-,\text{ and } 14n^0\ (28 - 14)$$
$$^{29}\text{Si has } 14p^+,\ 14e^-,\text{ and } 15n^0\ (29 - 14)$$
$$^{30}\text{Si has } 14p^+,\ 14e^-,\text{ and } 16n^0\ (30 - 14)$$

FOLLOW-UP PROBLEM 2.4 How many protons, neutrons, and electrons are in **(a)** $^{11}_{5}Q$? **(b)** $^{41}_{20}R$? **(c)** $^{131}_{53}X$? What element symbols do Q, R, and X represent?

The mass of an atom is measured *relative* to the mass of an atomic standard. The modern atomic mass standard is the carbon-12 atom. Its mass is defined as *exactly* 12 atomic mass units. Thus, the **atomic mass unit (amu)** is $\frac{1}{12}$ the mass of a carbon-12 atom. Based on this standard, the ^{1}H atom has a mass of 1.008 amu; in other words, a ^{12}C atom has almost 12 times the mass of an ^{1}H atom. We will continue to use the term *atomic mass unit* in the text, although the name of the unit has been changed to the **dalton (Da);** thus, one ^{12}C atom has a mass of 12 daltons (12 Da, or 12 amu). The atomic mass unit, which is a unit of relative mass, has an absolute mass of 1.66054×10^{-24} g.

Literature Blauch, David N.; Schuh, Merlyn D.; Carroll, Felix A. "Determination of the Natural Abundances of Krypton and Xenon Isotopes Using Mass Spectrometry: A Demonstration of Isotopes and the Basis of Atomic Mass," *J. Chem. Educ.* **2002,** *79,* 584.

The isotopic makeup of an element is determined by **mass spectrometry,** a method for measuring the relative masses and abundances of atomic-scale particles very precisely (see the Tools of the Laboratory essay). For example, using a mass spectrometer, we measure the mass ratio of ^{28}Si to ^{12}C as

$$\frac{\text{Mass of } ^{28}\text{Si atom}}{\text{Mass of } ^{12}\text{C standard}} = 2.331411$$

From this mass ratio, we find the **isotopic mass** of the ^{28}Si atom, the mass of the isotope relative to the mass of the standard carbon-12 isotope:

$$\begin{aligned}\text{Isotopic mass of } ^{28}\text{Si} &= \text{measured mass ratio} \times \text{mass of } ^{12}\text{C}\\ &= 2.331411 \times 12\text{ amu} = 27.97693\text{ amu}\end{aligned}$$

Along with the isotopic mass, the mass spectrometer gives the relative abundance (fraction) of each isotope in a sample of the element. For example, the percent abundance of ^{28}Si is 92.23%. Such measurements provide data for obtaining the **atomic mass** (also called *atomic weight*) of an element, the *average* of the masses of its naturally occurring isotopes weighted according to their abundances.

Literature Last, Arthur M.; Webb, Michael J. "Using Monetary Analogies to Teach Average Atomic Mass," *J. Chem. Educ.* **1993,** *70,* 234.

Literature Marshall, James L.; Marshall, Virginia R. "Rediscovery of the Elements: Ytterby Gruva (Ytterby Mine)," *J. Chem. Educ.* **2001,** *78,* 1343.

Literature Spindel, William; Ishida, Takanobu. "Isotope Separation," *J. Chem. Educ.* **1991,** *68,* 312.

Each naturally occurring isotope of an element contributes a certain portion to the atomic mass. For instance, as just noted, 92.23% of Si atoms are ^{28}Si. Using this percent abundance as a fraction and multiplying by the isotopic mass of ^{28}Si gives the portion of the atomic mass of Si contributed by ^{28}Si:

$$\text{Portion of Si atomic mass from } ^{28}\text{Si} = 27.97693\text{ amu} \times 0.9223 = 25.8031\text{ amu}$$
(retaining two additional significant figures)

Similar calculations give the portions contributed by ^{29}Si (28.976495 amu $\times$ 0.0467 = 1.3532 amu) and by ^{30}Si (29.973770 amu $\times$ 0.0310 = 0.9292 amu), and adding the three portions together (rounding to two decimal places at the end) gives the atomic mass of silicon:

$$\begin{aligned}\text{Atomic mass of Si} &= 25.8031\text{ amu} + 1.3532\text{ amu} + 0.9292\text{ amu}\\ &= 28.0855\text{ amu} = 28.09\text{ amu}\end{aligned}$$

Tools of the Laboratory

Mass Spectrometry

Mass spectrometry, the most powerful technique for measuring the mass and abundance of charged particles, emerged from electric and magnetic deflection studies on particles formed in cathode ray experiments. When a high-energy electron collides with an atom of neon-20, for example, one of the atom's electrons is knocked away and the resulting particle has one positive charge, Ne^+ (Figure B2.1). Thus, its mass/charge ratio (m/e) equals the mass divided by 1+. The m/e values are measured to identify the masses of different isotopes of an element.

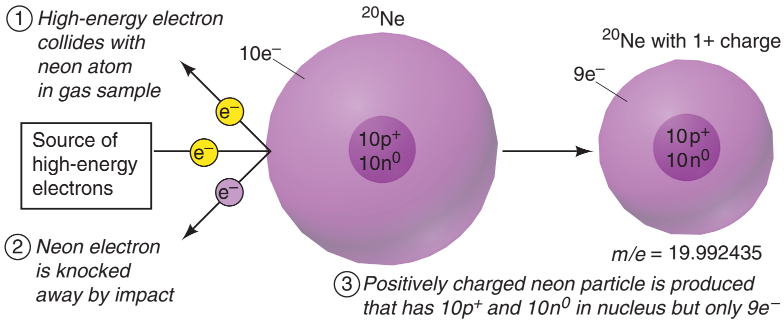
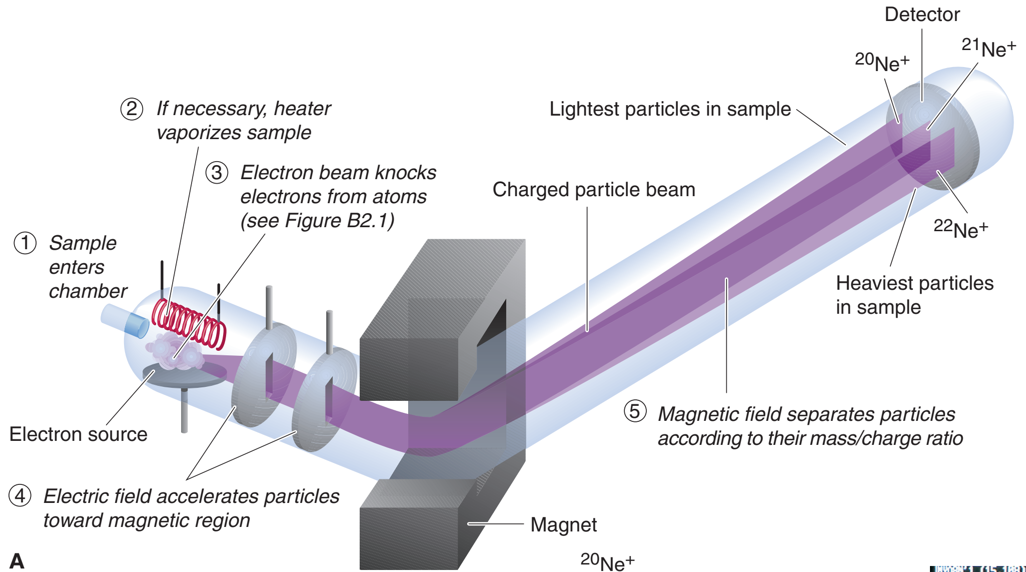
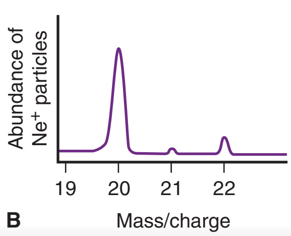


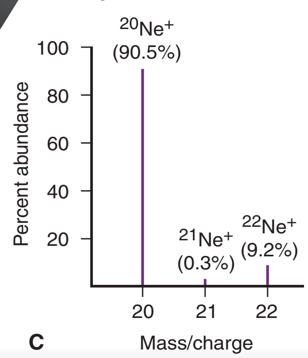
Figure B2.1 Formation of a positively charged neon (Ne) particle.

Figure B2.2, parts A–C, depicts the core of one type of mass spectrometer and the data it provides. The sample is introduced and vaporized (if liquid or solid), then bombarded by high-energy electrons to form positively charged particles. These are attracted toward a series of negatively charged plates with slits in them, and some particles pass through into an evacuated tube exposed to a magnetic field. As the particles zoom through this region, they are deflected (their paths are bent) according to their m/e: the lightest particles are deflected most and the heaviest particles least. At the end of the magnetic region, the particles strike a detector, which records their relative positions and abundances. For very precise work, such as determining isotopic masses and abundances, the instrument is calibrated with a substance of known amount and mass.

Mass spectrometry is also used in structural chemistry and separations science to measure the mass of virtually any atom, molecule, or molecular fragment. The technique is employed by biochemists determining protein structures (Figure B2.2, part D), materials scientists examining catalyst surfaces, forensic chemists analyzing criminal evidence, pharmaceutical chemists designing new drugs, industrial chemists investigating petroleum components, and many others. In fact, John B. Fenn and Koichi Tanaka shared part of the 2002 Nobel Prize in chemistry for developing methods to study proteins by mass spectrometry.







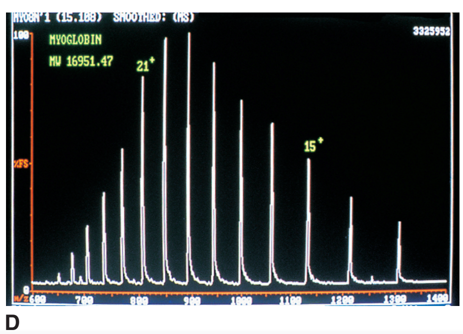


Figure B2.2 The mass spectrometer and its data. **A,** Charged particles are separated on the basis of their m/e values. Ne is the sample here. **B,** The data show the abundances of three Ne isotopes. **C,** The percent abundance of each isotope. **D,** The mass spectrum of a protein molecule. Each peak represents a molecular fragment.

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Note that atomic mass is an average value, and averages must be interpreted carefully. Although the average number of children in an American family in 1985 was 2.4, no family actually had 2.4 children; similarly, no individual silicon atom has a mass of 28.09 amu. But for most laboratory purposes, we consider a sample of silicon to consist of atoms with this average mass.

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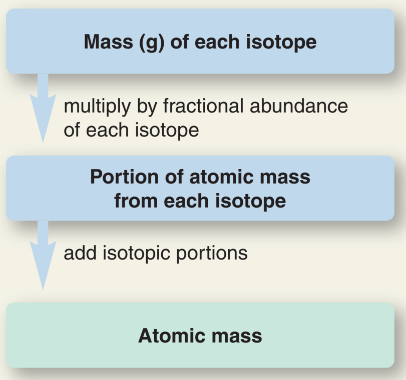
SAMPLE PROBLEM 2.5 Calculating the Atomic Mass of an Element

PROBLEM Silver (Ag; $Z = 47$) has 46 known isotopes, but only two occur naturally, ^{107}Ag and ^{109}Ag. Given the following mass spectrometric data, calculate the atomic mass of Ag:

Isotope	Mass (amu)	Abundance (%)
^{107}Ag	106.90509	51.84
^{109}Ag	108.90476	48.16



PLAN From the mass and abundance of the two Ag isotopes, we have to find the atomic mass of Ag (weighted average of the isotopic masses). We multiply each isotopic mass by its fractional abundance to find the portion of the atomic mass contributed by each isotope. The sum of the isotopic portions is the atomic mass.

SOLUTION Finding the portion of the atomic mass from each isotope:

$$\begin{aligned}\text{Portion of atomic mass from } ^{107}\text{Ag:} &= \text{isotopic mass} \times \text{fractional abundance}\\ &= 106.90509 \text{ amu} \times 0.5184 = 55.42 \text{ amu}\\ \text{Portion of atomic mass from } ^{109}\text{Ag:} &= 108.90476 \text{ amu} \times 0.4816 = 52.45 \text{ amu}\end{aligned}$$

Finding the atomic mass of silver:

$$\text{Atomic mass of Ag} = 55.42 \text{ amu} + 52.45 \text{ amu} = \boxed{107.87 \text{ amu}}$$

CHECK The individual portions seem right: $\sim$100 amu $\times$ 0.50 = 50 amu. The portions should be almost the same because the two isotopic abundances are almost the same. We rounded each portion to four significant figures because that is the number of significant figures in the abundance values. This is the correct atomic mass (to two decimal places), as shown in the list of elements (*inside front cover*).

FOLLOW-UP PROBLEM 2.5 Boron (B; $Z = 5$) has two naturally occurring isotopes. Find the percent abundances of ^{10}B and ^{11}B given the atomic mass of B = 10.81 amu, the isotopic mass of ^{10}B = 10.0129 amu, and the isotopic mass of ^{11}B = 11.0093 amu. (*Hint:* The sum of the fractional abundances is 1. If x = abundance of ^{10}B, then $1 - x$ = abundance of ^{11}B.)

Literature Reeves, Perry C.; Pamplin, Kim L. "A Strategy for Incorporating Hands-On GC-MS into the General Chemistry Lecture and Laboratory Courses," *J. Chem. Educ.* **2001,** *78,* 368.

Demo Ellis, Arthur B.; Adler, Edward A.; Juergens, Frederick H. "Dramatizing Isotopes: Deuterated Ice Cubes Sink," *J. Chem. Educ.* **1990,** *67,* 159.

A Modern Reassessment of the Atomic Theory

We began discussing the atomic basis of matter with Dalton's model, which proved inaccurate in several respects. What happens to a model whose postulates are found by later experiment to be incorrect? No model can predict every possible future observation, but a powerful model evolves and remains useful. Let's reexamine the atomic theory in light of what we know now:

1. *All matter is composed of atoms.* We now know that atoms *are* divisible and composed of smaller, subatomic particles (electrons, protons, and neutrons), but the atom is still the smallest body that *retains the unique identity* of an element.
2. *Atoms of one element cannot be converted into atoms of another element in a chemical reaction.* We now know that, in *nuclear* reactions, atoms of one element often change into atoms of another, but this *never* happens in a *chemical* reaction. ●
3. *All atoms of an element have the same number of protons and electrons, which determines the chemical behavior of the element.* We now know that isotopes of an element differ in the number of neutrons, and thus in mass number, but a sample of the element is treated as though its atoms have an *average* mass.
4. *Compounds are formed by the chemical combination of two or more elements in specific ratios.* We now know that a few compounds can have slight variations in their atom ratios, but this postulate remains essentially unchanged.

● **The Heresy of Radioactive "Transmutation"** In 1902, Rutherford performed a series of experiments with radioactive elements that shocked the scientific world. When a radioactive atom of thorium ($Z = 90$) emits an α particle ($Z = 2$), it becomes an atom of radium ($Z = 88$), which then emits another α particle and becomes an atom of radon ($Z = 86$). He proposed that when an atom emits an α particle, it turns into a different atom—one element changes into another! Many viewed this conclusion as a return to alchemy, and, as with Thomson's discovery that atoms contain smaller particles, Rutherford's findings fell on disbelieving ears.

Even today, our picture of the atom is being revised. Although we are confident about the distribution of electrons within the atom (Chapters 7 and 8), the interactions among protons and neutrons within the nucleus are still on the frontier of discovery (Chapter 24).

Section Summary

An atom has a central nucleus, which contains positively charged protons and uncharged neutrons and is surrounded by negatively charged electrons. An atom is neutral because the number of electrons equals the number of protons. • An atom is represented by the notation ${}^{A}_{Z}X$, in which Z is the atomic number (number of protons), A the mass number (sum of protons and neutrons), and X the atomic symbol. • An element occurs naturally as a mixture of isotopes, atoms with the same number of protons but different numbers of neutrons. Each isotope has a mass relative to the ${}^{12}C$ mass standard. • The atomic mass of an element is the average of its isotopic masses weighted according to their natural abundances and is determined by modern instruments, especially the mass spectrometer.

2.6 ELEMENTS: A FIRST LOOK AT THE PERIODIC TABLE

At the end of the 18th century, Lavoisier compiled a list of the 23 elements known at that time; by 1870, 65 were known; by 1925, 88; today, there are 116 and still counting! These elements combine to form millions of compounds, so we clearly need some way to organize what we know about their behavior. By the mid-19th century, enormous amounts of information concerning reactions, properties, and atomic masses of the elements had been accumulated. Several researchers noted recurring, or *periodic,* patterns of behavior and proposed schemes to organize the elements according to some fundamental property.

In 1871, the Russian chemist Dmitri Mendeleev (1836–1907) published the most successful of these organizing schemes as a table of the elements listed by increasing atomic mass and arranged so that elements with similar chemical properties fell in the same column. The modern **periodic table of the elements,** based on Mendeleev's earlier version (but arranged by atomic number, not mass), is one of the great classifying schemes in science and is now an indispensable tool to chemists. Throughout your study of chemistry, the periodic table will guide you through an otherwise dizzying amount of chemical and physical behavior.

Literature Talanquer, Vicente. "Recreating a Periodic Table: A Tool for Developing Pedagogical Content Knowledge," *Chem. Educator* 10 **2005,** *2,* 95–99.

Literature Dopke, Nancy Carter; Treichel, Paul M., Jr.; Vestling, Martha M. "Significant Figures, the Periodic Table, and Mass Spectrometry: The Challenge of Large Biomolecules," *J. Chem. Educ.* **2000,** *77,* 1065.

Literature Dutch, Steven I. "Periodic Tables of Elemental Abundance," *J. Chem. Educ.* **1999,** *76,* 356.

Organization of the Periodic Table A modern version of the periodic table appears in Figure 2.9 on the next page and inside the front cover. It is formatted as follows:

1. Each element has a box that contains its atomic number, atomic symbol, and atomic mass. The boxes lie in order of *increasing atomic number* (number of protons) as you move from left to right.
2. The boxes are arranged into a grid of **periods** (horizontal rows) and **groups** (vertical columns). Each period has a number from 1 to 7. Each group has a number from 1 to 8 *and* either the letter A or B. A new system, with group numbers from 1 to 18 but no letters, appears in parentheses under the number-letter designations. (Most chemists still use the number-letter system, so the text retains it, but shows the new numbering system in parentheses.)
3. The eight A groups (two on the left and six on the right) contain the *main-group,* or *representative, elements.* The ten B groups, located between Groups 2A(2) and 3A(13), contain the *transition elements.* Two horizontal series of *inner transition elements,* the lanthanides and the actinides, fit *between* the elements in Group 3B(3) and Group 4B(4) and are usually placed below the main body of the table.

Literature He, Fu-cheng; Li, Xiang-yuan. "The Periodic Building of the Elements: Can a Periodic Table Be Transformed into a Stereo One?" *J. Chem. Educ.* **1997,** *74,* 792.

Metals (main-group)
Metals (transition)
Metals (inner transition)
Metalloids
Nonmetals

MAIN–GROUP ELEMENTS | TRANSITION ELEMENTS | MAIN–GROUP ELEMENTS

Period	1A (1)	2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
1	1 **H** 1.008																	2 **He** 4.003
2	3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.88	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.41	31 **Ga** 69.72	32 **Ge** 72.61	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
6	55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.9	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (263)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (267)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (281)	111 **Rg** (272)	112 (285)	113 (284)	114 (289)	115 (288)	116 (292)		

INNER TRANSITION ELEMENTS

Period															
6	Lanthanides	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0	71 **Lu** 175.0
7	Actinides	90 **Th** 232.0	91 **Pa** (231)	92 **U** 238.0	93 **Np** (237)	94 **Pu** (242)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)

Figure 2.9 The modern periodic table. The table consists of element boxes arranged by *increasing* atomic number into groups (vertical columns) and periods (horizontal rows). Each box contains the atomic number, atomic symbol, and atomic mass. (A mass in parentheses is the mass number of the most stable isotope of that element.) The periods are numbered 1 to 7. The groups (sometimes called *families*) have a number-letter designation and a new group number in parentheses. The A groups are the main-group elements; the B groups are the transition elements. Two series of inner transition elements are placed below the main body of the table but actually fit between the elements indicated. Metals lie below and to the left of the thick "staircase" line [top of 3A(13) to bottom of 6A(16) in Period 6] and include main-group metals (*purple-blue*), transition elements (*blue*), and inner transition elements (*gray-blue*). Nonmetals (*yellow*) lie to the right of the line. Metalloids (*green*) lie along the line. We discuss the placement of hydrogen in Chapter 14. As of mid-2007, elements 112–116 had not been named.

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At this point in the text, the clearest distinction among the elements is their classification as metals, nonmetals, or metalloids. The "staircase" line that runs from the top of Group 3A(13) to the bottom of Group 6A(16) in Period 6 is a dividing line for this classification. The **metals** (three shades of blue) appear in the large lower-left portion of the table. About three-quarters of the elements are metals, including many main-group elements and all the transition and inner transition elements. They are generally shiny solids at room temperature (mercury is the only liquid) that conduct heat and electricity well and can be tooled into sheets (malleable) and wires (ductile). The **nonmetals** (yellow) appear in the small upper-right portion of the table. They are generally gases or dull, brittle solids at room temperature (bromine is the only liquid) and conduct heat and electricity poorly. Along the staircase line lie the **metalloids** (green; also called **semimetals**), elements that have properties between those of metals and nonmetals. Several

Literature Jensen, William B. "Why Helium Ends in "-ium"," *J. Chem. Educ.* **2004,** *81,* 944.

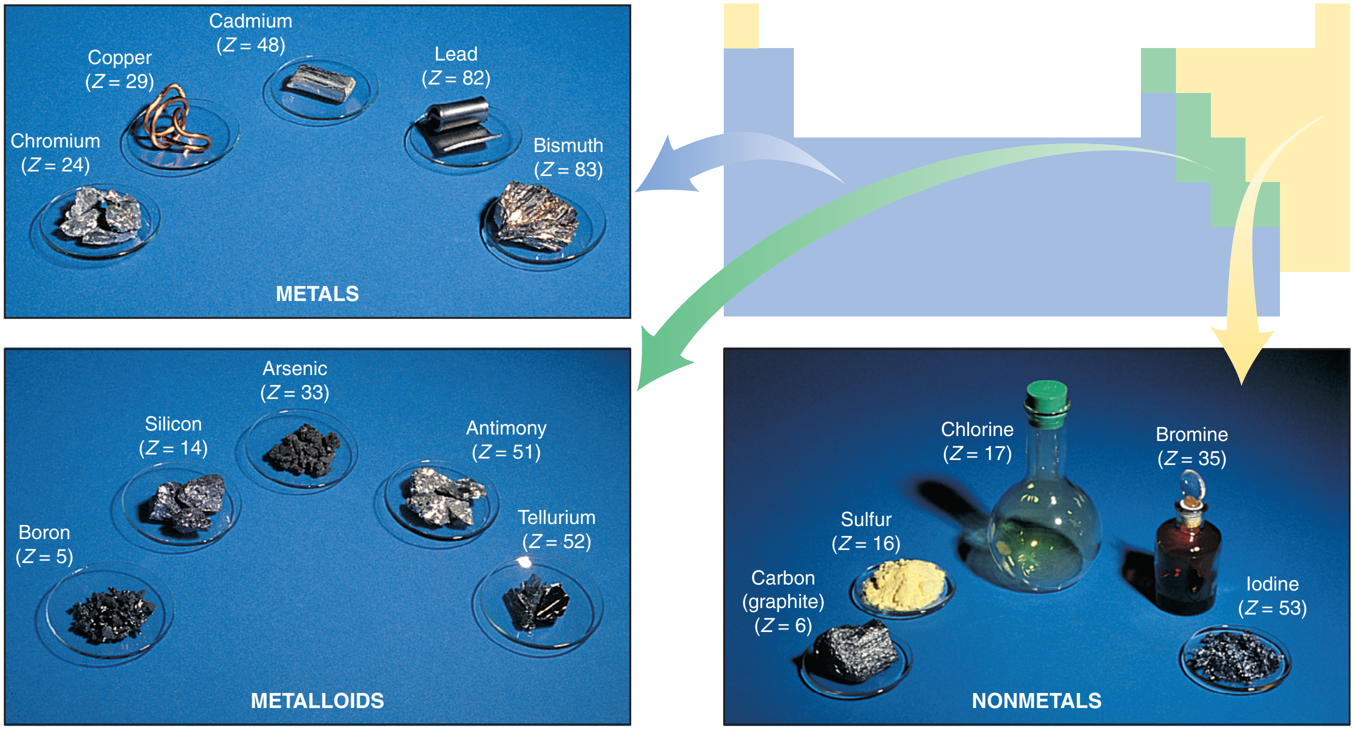

Figure 2.10 **Some metals, metalloids, and nonmetals.**

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metalloids, such as silicon (Si) and germanium (Ge), play major roles in modern electronics. Figure 2.10 shows examples of these three classes of elements.

Two of the major branches of chemistry have traditionally been defined by the elements that each studies. *Organic chemistry* studies the compounds of carbon, specifically those that contain hydrogen and often oxygen, nitrogen, and a few other elements. This branch is concerned with fuels, drugs, dyes, polymers, and the like. *Inorganic chemistry,* on the other hand, focuses mainly on the compounds of all the other elements. It is concerned with catalysts, electronic materials, metal alloys, mineral salts, and the like. With the explosive growth in biomedical and materials research, the line between these branches has all but disappeared.

Literature Nelson, P. G. "Important Elements," *J. Chem. Educ.* **1991,** *68,* 732.

It is important to learn some of the group (family) names. Group 1A(1), except for hydrogen, consists of the *alkali metals,* and Group 2A(2) consists of the *alkaline earth metals.* Both groups consist of highly reactive elements. The *halogens,* Group 7A(17), are highly reactive nonmetals, whereas the *noble gases,* Group 8A(18), are relatively unreactive nonmetals. Other main groups [3A(13) to 6A(16)] are often named for the first element in the group; for example, Group 6A is the *oxygen family.*

Literature Ringnes, Vivi. "Origin of the Names of Chemical Elements," *J. Chem. Educ.* **1989,** *66,* 731.

Literature Cronyn, Marshall W. "The Proper Place for Hydrogen in the Periodic Table," *J. Chem. Educ.* **2003,** *80,* 947.

A key point that we return to many times is that, in general, *elements in a group have* ***similar*** *chemical properties and elements in a period have* ***different*** *chemical properties.* We begin applying the organizing power of the periodic table in the next section, where we discuss how elements combine to form compounds.

Section Summary

In the periodic table, the elements are arranged by atomic number into horizontal periods and vertical groups. • Because of the periodic recurrence of certain key properties, elements within a group have similar behavior, whereas elements in a period have dissimilar behavior. • Nonmetals appear in the upper-right portion of the table, metalloids lie along a staircase line, and metals fill the rest of the table.

2.7 COMPOUNDS: INTRODUCTION TO BONDING

The overwhelming majority of elements occur in chemical combination with other elements. In fact, only a few elements occur free in nature. The noble gases—helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn)—occur in air as separate atoms. In addition to occurring in compounds, oxygen (O), nitrogen (N), and sulfur (S) occur in the most common elemental form as the molecules O_2, N_2, and S_8, and carbon (C) occurs in vast, nearly pure deposits of coal. Some of the metals, such as copper (Cu), silver (Ag), gold (Au), and platinum (Pt), may also occur uncombined with other elements. But these few exceptions reinforce the general rule that elements occur combined in compounds.

It is the electrons of the atoms of interacting elements that are involved in compound formation. Elements combine in two general ways:

1. *Transferring electrons* from the atoms of one element to those of another to form **ionic compounds**
2. *Sharing electrons* between atoms of different elements to form **covalent compounds**

These processes generate **chemical bonds,** the forces that hold the atoms of elements together in a compound. We'll introduce compound formation next and have much more to say about it in later chapters.

Animation: Formation of an Ionic Compound
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Chapter 2

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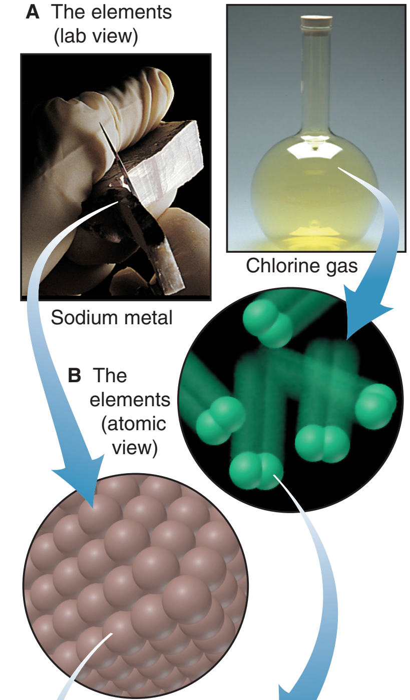
The Formation of Ionic Compounds

Ionic compounds are composed of **ions,** charged particles that form when an atom (or small group of atoms) gains or loses one or more electrons. The simplest type of ionic compound is a **binary ionic compound,** one composed of just two elements. It typically forms *when a metal reacts with a nonmetal.* Each metal atom loses a certain number of its electrons and becomes a **cation,** a positively charged ion. The nonmetal atoms gain the electrons lost by the metal atoms and become **anions,** negatively charged ions. In effect, the metal atoms *transfer electrons* to the nonmetal atoms. The resulting cations and anions attract each other through electrostatic forces and form the ionic compound. *All binary ionic compounds are solids.* A cation or anion derived from a single atom is called a **monatomic ion;** we'll discuss polyatomic ions, those derived from a small group of atoms, later.

The formation of the binary ionic compound sodium chloride, common table salt, is depicted in Figure 2.11, from the elements through the atomic-scale



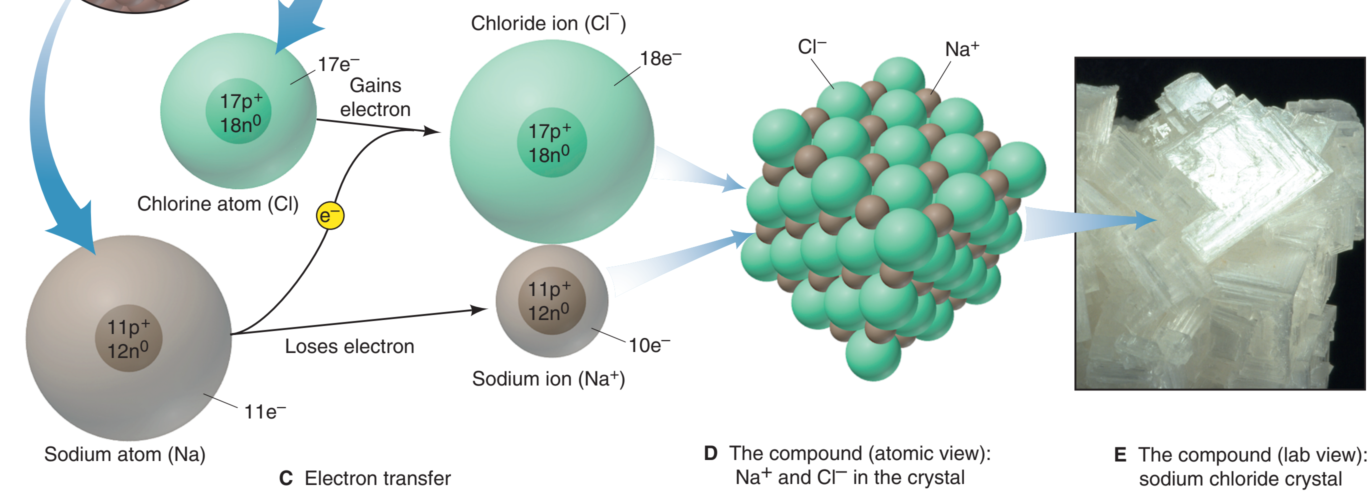


Figure 2.11 The formation of an ionic compound. A, The two elements as seen in the laboratory. **B,** The elements on the atomic scale. **C,** The neutral sodium atom loses one electron to become a sodium cation (Na^+), and the chlorine atom gains one electron to become a chloride anion (Cl^-). (Note that when atoms lose electrons, they become ions that are smaller, and when they gain electrons, they become ions that are larger.) **D,** Na^+ and Cl^- ions attract each other and form a regular three-dimensional array. **E,** This arrangement of the ions is reflected in the structure of crystalline NaCl, which occurs naturally as the mineral halite, hence the name *halogens* for the Group 7A(17) elements.

electron transfer to the compound. In the electron transfer, a sodium atom, which is neutral because it has the same number of protons as electrons, *loses* one electron and forms a sodium cation, Na^+. (The charge on the ion is written as a *right superscript*.) A chlorine atom *gains* the electron and becomes a chloride anion, Cl^-. (The name change from the nonmetal atom to the ion is discussed in the next section.) Even the tiniest visible grain of table salt contains an enormous number of sodium and chloride ions. The oppositely charged ions (Na^+ and Cl^-) attract each other, and the similarly charged ions (Na^+ and Na^+, or Cl^- and Cl^-) repel each other. The resulting solid aggregation is a regular array of alternating Na^+ and Cl^- ions that extends in all three dimensions.

The strength of the ionic bonding depends to a great extent on the net strength of these attractions and repulsions and is described by *Coulomb's law,* which can be expressed as follows: *the energy of attraction (or repulsion) between two particles is directly proportional to the product of the charges and inversely proportional to the distance between them.*

$$\text{Energy} \propto \frac{\text{charge 1} \times \text{charge 2}}{\text{distance}}$$

In other words, ions with higher charges attract (or repel) each other more strongly than ions with lower charges. Likewise, smaller ions attract (or repel) each other more strongly than larger ions, because their charges are closer together. These effects are summarized in Figure 2.12.

Ionic compounds are neutral; that is, they possess no net charge. For this to occur, they must contain equal numbers of positive and negative *charges*—not necessarily equal numbers of positive and negative *ions*. Because Na^+ and Cl^- each bear a unit charge (1+ or 1−), equal numbers of these ions are present in sodium chloride; but in sodium oxide, for example, there are two Na^+ ions present to balance the 2− charge of each oxide ion, O^{2-}.

Can we predict the number of electrons a given atom will lose or gain when it forms an ion? In the formation of sodium chloride, for example, why does each sodium atom give up only 1 of its 11 electrons? Why doesn't each chlorine atom gain two electrons, instead of just one? For A-group elements, the periodic table provides an answer. We generally find that metals lose electrons and nonmetals gain electrons to *form ions with the same number of electrons as in an atom of the nearest noble gas* [Group 8A(18)]. Noble gases have a stability (low reactivity) that is related to their number (and arrangement) of electrons. A sodium atom ($11e^-$) can attain the stability of neon ($10e^-$), the nearest noble gas, by losing one electron. Similarly, by gaining one electron, a chlorine atom ($17e^-$) attains the stability of argon ($18e^-$), its nearest noble gas. Thus, when an element located near a noble gas forms a monatomic ion, *it gains or loses enough electrons to attain the same number as that noble gas.* Specifically, the elements that are in Group 1A(1) lose one electron, those in Group 2A(2) lose two, and aluminum in Group 3A(13) loses three; the elements in Group 7A(17) gain one electron, oxygen and sulfur in Group 6A(16) gain two, and nitrogen in Group 5A(15) gains three.

With the periodic table printed on a two-dimensional surface, as in Figure 2.9, it is easy to get the false impression that the elements in Group 7A(17) are "closer" to the noble gases than the elements in Group 1A(1). Actually, both groups are only one electron away from having the same number of electrons as the noble gases. To make this point, Figure 2.13 shows a modified periodic table of monatomic ions that is cut and rejoined as a cylinder. Now you can see that fluorine (F; $Z = 9$) has one electron fewer than the noble gas neon (Ne; $Z = 10$) and sodium (Na; $Z = 11$) has one electron more; thus, they form the F^- and Na^+ ions. Similarly, oxygen (O; $Z = 8$) gains two electrons and magnesium (Mg; $Z = 12$) loses two to form the O^{2-} and Mg^{2+} ions and attain the same number of electrons as neon.

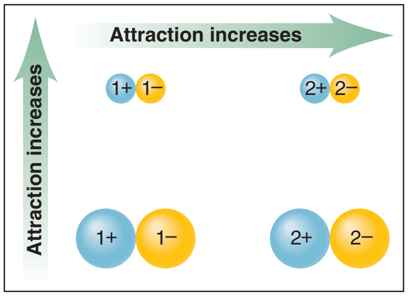


Figure 2.12 Factors that influence the strength of ionic bonding. For ions of a given size, strength of attraction *(arrows)* increases with higher ionic charge *(left to right).* For ions of a given charge, strength of attraction increases with smaller ionic size *(bottom to top).*

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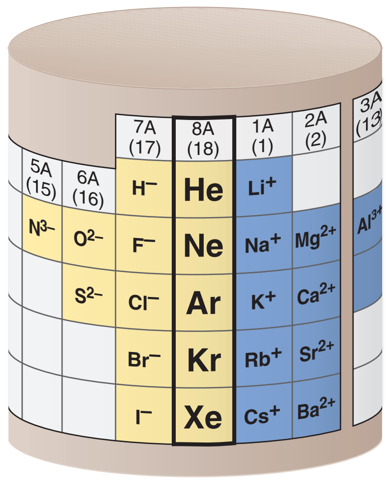

Figure 2.13 The relationship between ions formed and the nearest noble gas. This periodic table was redrawn to show the positions of other nonmetals *(yellow)* and metals *(blue)* relative to the noble gases and to show the ions these elements form. The ionic charge equals the number of electrons lost (+) or gained (−) to attain the same number of electrons as the nearest noble gas. Species in the same row have the same number of electrons. For example, H^-, He, and Li^+ all have two electrons. [Note that H is shown here in Group 7A(17).]

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SAMPLE PROBLEM 2.6 Predicting the Ion an Element Forms

PROBLEM What monatomic ions do the following elements form?
(a) Iodine (Z = 53) **(b)** Calcium (Z = 20) **(c)** Aluminum (Z = 13)

PLAN We use the given Z value to find the element in the periodic table and see where its group lies relative to the noble gases. Elements in Groups 1A, 2A, and 3A *lose* electrons to attain the same number as the nearest noble gas and become positive ions; those in Groups 5A, 6A, and 7A *gain* electrons and become negative ions.

SOLUTION **(a)** I^- Iodine ($_{53}I$) is a nonmetal in Group 7A(17), one of the halogens. Like any member of this group, it gains 1 electron to have the same number as the nearest Group 8A(18) member, in this case $_{54}Xe$.
(b) Ca^{2+} Calcium ($_{20}Ca$) is a member of Group 2A(2), the alkaline earth metals. Like any Group 2A member, it loses 2 electrons to attain the same number as the nearest noble gas, in this case, $_{18}Ar$.
(c) Al^{3+} Aluminum ($_{13}Al$) is a metal in the boron family [Group 3A(13)] and thus loses 3 electrons to attain the same number as its nearest noble gas, $_{10}Ne$.

FOLLOW-UP PROBLEM 2.6 What monatomic ion does each of the following elements form: **(a)** $_{16}S$; **(b)** $_{37}Rb$; **(c)** $_{56}Ba$?

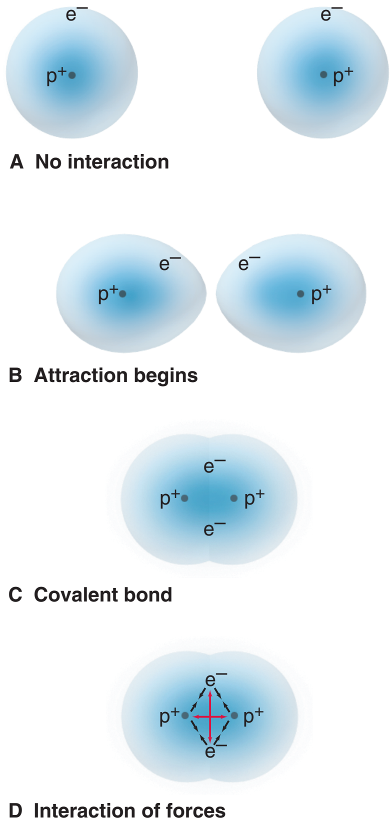

Figure 2.14 Formation of a covalent bond between two H atoms. A, The distance is too great for the atoms to affect each other. **B,** As the distance decreases, the nucleus of each atom begins to attract the electron of the other. **C,** The covalent bond forms when the two nuclei mutually attract the pair of electrons at some optimum distance. **D,** The H_2 molecule is more stable than the separate atoms because the attractive forces *(black arrows)* between each nucleus and the two electrons are greater than the repulsive forces *(red arrows)* between the electrons and between the nuclei.

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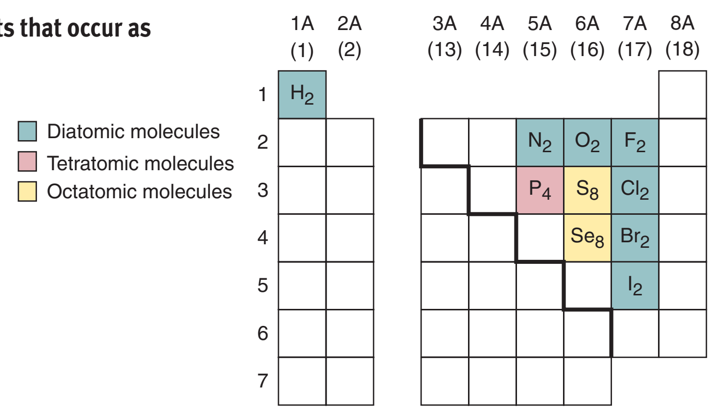
The Formation of Covalent Compounds

Covalent compounds form when elements share electrons, which usually occurs between nonmetals. Even though relatively few nonmetals exist, they interact in many combinations to form a very large number of covalent compounds.

The simplest case of electron sharing occurs not in a compound but between two hydrogen atoms (H; $Z = 1$). Imagine two separated H atoms approaching each other, as in Figure 2.14. As they get closer, the nucleus of each atom attracts the electron of the other atom more and more strongly, and the separated atoms begin to interpenetrate each other. At some optimum distance between the nuclei, the two atoms form a **covalent bond,** a pair of electrons mutually attracted by the two nuclei. The result is a hydrogen molecule, in which each electron no longer "belongs" to a particular H atom: the two electrons are *shared* by the two nuclei. Repulsions between the nuclei and between the electrons also occur, but the net attraction is greater than the net repulsion. (We discuss the properties of covalent bonds in great detail in Chapter 9.)

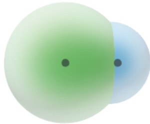
A sample of hydrogen gas consists of these diatomic molecules (H_2)—pairs of atoms that are chemically bound and behave as an independent unit—*not* separate H atoms. Other nonmetals that exist as diatomic molecules at room temperature are nitrogen (N_2), oxygen (O_2), and the halogens [fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2)]. Phosphorus exists as tetratomic molecules (P_4), and sulfur and selenium as octatomic molecules (S_8 and Se_8) (Figure 2.15). At room temperature, covalent substances may be gases, liquids, or solids.

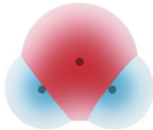
Figure 2.15 Elements that occur as molecules.



Atoms of different elements share electrons to form the molecules of a covalent compound. A sample of hydrogen fluoride, for example, consists of molecules in which one H atom forms a covalent bond with one F atom; water consists of molecules in which one O atom forms covalent bonds with two H atoms:

Hydrogen fluoride, HF





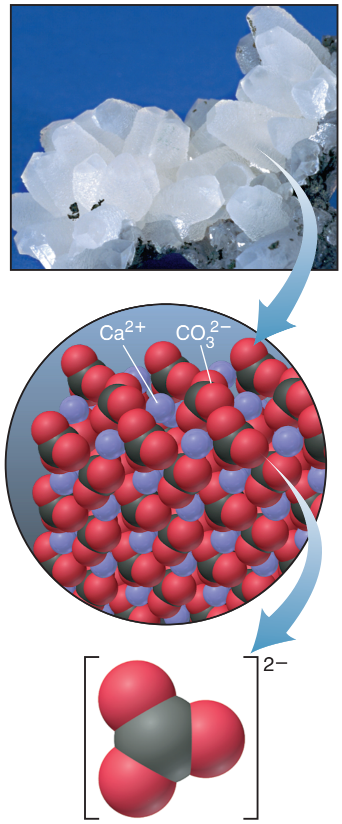
Water, H_2O

(As you'll see in Chapter 9, covalent bonding provides another way for atoms to attain the same number of electrons as the nearest noble gas.)

Distinguishing the Entities in Covalent and Ionic Substances There is a key distinction between the chemical entities in covalent substances and ionic substances. *Most covalent substances consist of molecules.* A cup of water, for example, consists of individual water molecules lying near each other. In contrast, under ordinary conditions, *no molecules exist in a sample of an ionic compound.* A piece of sodium chloride, for example, is a continuous array of oppositely charged sodium and chloride ions, *not* a collection of individual "sodium chloride molecules."

Another key distinction exists between the nature of the particles attracting each other. Covalent bonding involves the mutual attraction between two (positively charged) nuclei and the two (negatively charged) electrons that reside between them. Ionic bonding involves the mutual attraction between positive and negative ions.

Polyatomic Ions: Covalent Bonds Within Ions Many ionic compounds contain **polyatomic ions,** which consist of two or more atoms bonded *covalently* and have a net positive or negative charge. For example, the ionic compound calcium carbonate is an array of polyatomic carbonate anions and monatomic calcium cations attracted to each other. The carbonate ion consists of a carbon atom covalently bonded to three oxygen atoms, and two additional electrons give the ion its 2− charge (Figure 2.16). In many reactions, a polyatomic ion stays together as a unit.



Carbonate ion
CO_3^{2-}

Figure 2.16 A polyatomic ion. Calcium carbonate is a three-dimensional array of monatomic calcium cations *(purple spheres)* and polyatomic carbonate anions. As the bottom structure shows, each carbonate ion consists of four covalently bonded atoms.

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Literature Marshall, James L. "A Living Periodic Table," *J. Chem. Educ.* **2000,** *77,* 979.

Biological application

The Elements of Life

About one-quarter of all the elements have known roles in organisms. As you can see in Figure 2.17, metals, nonmetals, and metalloids are among these essential elements. But, except for some diatomic oxygen and nitrogen molecules inhaled into the lungs, none of the elements in organisms occurs in pure form; rather, they appear in compounds or as ions in solution.

Building-block elements
Major minerals
Trace elements

	1A (1)	2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
1	H																	
2													B	C	N	O	F	
3	Na	Mg												Si	P	S	Cl	
4	K	Ca			V	Cr	Mn	Fe	Co	Ni	Cu	Zn			As	Se		
5						Mo								Sn			I	

Figure 2.17 A biological periodic table. The building-block elements and major minerals are required by all organisms. Most organisms, including humans, require the trace elements as well. Many other elements (not shown) are found in organisms but have no known role.

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The elements of life are often classified by the amount present in organisms. The four nonmetals carbon (C), oxygen (O), hydrogen (H), and nitrogen (N) are the *building-block elements* because they make up the major portion of biological molecules. Over 99% of the atoms in organisms are C, O, H, and N; in humans, they account for over 96% by mass of body weight. The nonmetals O and H make up the water in organisms, of course, and together with C occur in all four major classes of biological molecules—carbohydrates, fats, proteins, and nucleic acids. All proteins and nucleic acids also contain N.

The seven *major minerals* (or *macronutrients*) range from around 2% by mass for calcium (Ca) to around 0.14% by mass for chlorine (Cl). The alkali metals sodium and potassium and the halogen chlorine are dissolved in cell fluids as the ions Na^+, K^+, and Cl^-. The alkaline earth metals magnesium and calcium occur as Mg^{2+} and Ca^{2+}, most often bound to proteins or, in the case of calcium, in bones and teeth. Sulfur (S) occurs mostly in proteins, but phosphorus (P) also occurs in nucleic acids, many fats, and sugars, and as part of a polyatomic ion in bone and cell fluids.

The *trace elements* (or *micronutrients*) are present in much lower amounts, with iron (Fe) the most abundant at only 0.005% by mass. Most of them are associated with protein functions. We will look more closely at the trace elements in Chapter 23.

Section Summary

Although a few elements occur uncombined in nature, the great majority exist in compounds. • Ionic compounds form when a metal *transfers* electrons to a nonmetal, and the resulting positive and negative ions attract each other to form a three-dimensional array. In many cases, metal atoms lose and nonmetal atoms gain enough electrons to attain the same number of electrons as in atoms of the nearest noble gas. • Covalent compounds form when elements, usually nonmetals, *share* electrons. Each covalent bond is an electron pair mutually attracted by two atomic nuclei. • Monatomic ions are derived from single atoms. Polyatomic ions consist of two or more covalently bonded atoms that have a net positive or negative charge due to a deficit or excess of electrons. • The elements in organisms are found as ions or bonded in large biomolecules. Four building-block elements (C, O, H, N) form these compounds, seven other elements (major minerals, or macronutrients) are also common, and many others (trace elements, or micronutrients) occur in tiny amounts and play specific roles.

2.8 COMPOUNDS: FORMULAS, NAMES, AND MASSES

Literature Wirtz, Michael C.; Kaufmann, Joan; Hawley, Gary. "Nomenclature Made Practical: Student Discovery of the Nomenclature Rules," *J. Chem. Educ.* **2006,** *83,* 595.

Literature Chimeno, Joseph. "How to Make Learning Chemical Nomenclature Fun, Exciting, and Palatable," *J. Chem. Educ.* **2000,** *77,* 144.

Literature Lind, Gerhard. "Teaching Inorganic Nomenclature: A Systematic Approach," *J. Chem. Educ.* **1992,** *69,* 613.

Names and formulas of compounds form the vocabulary of the chemical language. In this discussion, you'll learn the names and formulas of ionic and simple covalent compounds and how to calculate the mass of a unit of a compound from its formula.

Types of Chemical Formulas

In a **chemical formula,** element symbols and numerical subscripts show the type and number of each atom present in the smallest unit of the substance. There are several types of chemical formulas for a compound:

1. The **empirical formula** shows the *relative* number of atoms of each element in the compound. It is the simplest type of formula and is derived from the masses of the component elements. For example, in hydrogen peroxide, there is 1 part by mass of hydrogen for every 16 parts by mass of oxygen. Therefore, the empirical formula of hydrogen peroxide is HO: one H atom for every O atom.

2. The **molecular formula** shows the *actual* number of atoms of each element in a molecule of the compound. The molecular formula of hydrogen peroxide is H_2O_2; there are two H atoms and two O atoms in each molecule.
3. A **structural formula** shows the number of atoms and *the bonds between them;* that is, the relative placement and connections of atoms in the molecule. The structural formula of hydrogen peroxide is H—O—O—H; each H is bonded to an O, and the O's are bonded to each other.

Some Advice about Learning Names and Formulas

Perhaps in the future, systematic names for compounds will be used by everyone. However, many reference books, chemical supply catalogs, and practicing chemists still use many common (trivial) names, so you should learn them as well.

Here are some points to note about ion formulas:

- Members of a periodic table group have the same ionic charge; for example, Li, Na, and K are all in Group 1A and all have a 1+ charge.
- For A-group cations, ion charge = group number: for example, Na^+ is in Group 1A, Ba^{2+} in Group 2A. (Exceptions in Figure 2.18 are Sn^{2+} and Pb^{2+}.)
- For anions, ion charge = group number minus 8: for example, S is in Group 6A ($6 - 8 = -2$), so the ion is S^{2-}.

Here are some suggestions about how to learn names and formulas:

1. Memorize the A-group monatomic ions of Table 2.3 (all except Ag^+, Zn^{2+}, and Cd^{2+}) according to their positions in Figure 2.18. These ions have the same number of electrons as an atom of the nearest noble gas.
2. Consult Table 2.4 (page 67) and Figure 2.18 for some metals that form two different monatomic ions.
3. Divide the tables of names and charges into smaller batches, and learn a batch each day. Try flash cards, with the name on one side and the ion formula on the other. The most common ions are shown in **boldface** in Tables 2.3, 2.4, and 2.5, so you can focus on learning them first.

Table 2.3 Common Monatomic Ions*

Charge	Formula	Name
Cations		
1+	H^+	hydrogen
	Li^+	**lithium**
	Na^+	**sodium**
	K^+	**potassium**
	Cs^+	cesium
	Ag^+	**silver**
2+	**Mg^{2+}**	**magnesium**
	Ca^{2+}	**calcium**
	Sr^{2+}	strontium
	Ba^{2+}	**barium**
	Zn^{2+}	**zinc**
	Cd^{2+}	cadmium
3+	**Al^{3+}**	**aluminum**
Anions		
1−	H^-	hydride
	F^-	**fluoride**
	Cl^-	**chloride**
	Br^-	**bromide**
	I^-	**iodide**
2−	**O^{2-}**	**oxide**
	S^{2-}	**sulfide**
3−	N^{3-}	nitride

*Listed by charge; those in **boldface** are most common.

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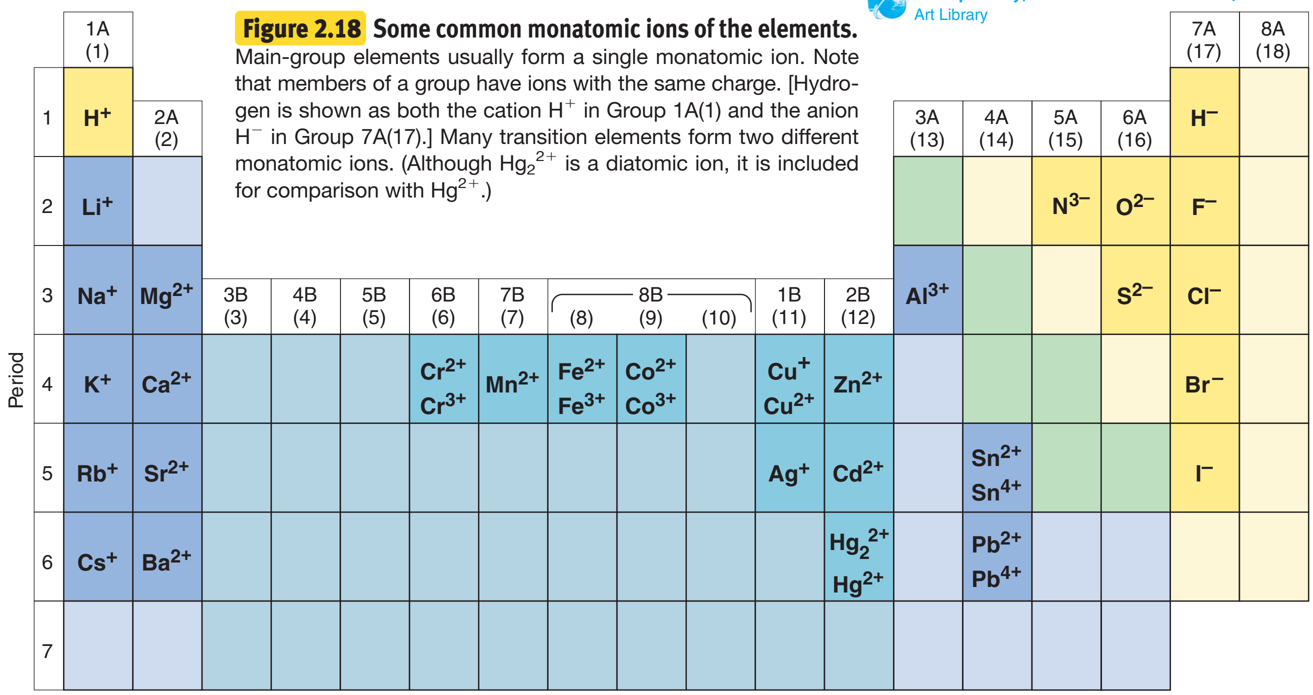

Figure 2.18 Some common monatomic ions of the elements. Main-group elements usually form a single monatomic ion. Note that members of a group have ions with the same charge. [Hydrogen is shown as both the cation H^+ in Group 1A(1) and the anion H^- in Group 7A(17).] Many transition elements form two different monatomic ions. (Although Hg_2^{2+} is a diatomic ion, it is included for comparison with Hg^{2+}.)

Names and Formulas of Ionic Compounds

All ionic compound names give the positive ion (cation) first and the negative ion (anion) second.

Compounds Formed from Monatomic Ions Let's first consider how to name binary ionic compounds, those composed of ions of two elements.

- *The name of the cation is the same as the name of the metal.* Many metal names end in *-ium*.
- *The name of the anion takes the root of the nonmetal name and adds the suffix -ide.*

For example, the anion formed from brom*ine* is named brom*ide* (brom+ide). Therefore, the compound formed from the metal calcium and the nonmetal bromine is named *calcium bromide*.

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SAMPLE PROBLEM 2.7 Naming Binary Ionic Compounds

PROBLEM Name the ionic compound formed from the following pairs of elements:

(a) Magnesium and nitrogen **(b)** Iodine and cadmium
(c) Strontium and fluorine **(d)** Sulfur and cesium

PLAN The key to naming a binary ionic compound is to recognize which element is the metal and which is the nonmetal. When in doubt, check the periodic table. We place the cation name first, add the suffix *-ide* to the nonmetal root, and place the anion name last.

SOLUTION **(a)** Magnesium is the metal; *nitr-* is the nonmetal root: magnesium nitride

(b) Cadmium is the metal; *iod-* is the nonmetal root: cadmium iodide

(c) Strontium is the metal; *fluor-* is the nonmetal root: strontium fluoride (Note the spelling is fl*uo*ride, not fl*ou*ride.)

(d) Cesium is the metal; *sulf-* is the nonmetal root: cesium sulfide

FOLLOW-UP PROBLEM 2.7 For the following ionic compounds, give the name and periodic table group number of each of the elements present: **(a)** zinc oxide; **(b)** silver bromide; **(c)** lithium chloride; **(d)** aluminum sulfide.

Ionic compounds are arrays of oppositely charged ions rather than separate molecular units. Therefore, we write a formula for the **formula unit,** which gives the *relative* numbers of cations and anions in the compound. Thus, ionic compounds generally have only empirical formulas.* The compound has zero net charge, so the positive charges of the cations must balance the negative charges of the anions. For example, calcium bromide is composed of Ca^{2+} ions and Br^- ions; therefore, two Br^- balance each Ca^{2+}. The formula is $CaBr_2$, not Ca_2Br. In this and all other formulas,

- The subscript refers to the element *preceding* it.
- The *subscript 1 is understood* from the presence of the element symbol alone (that is, we do not write Ca_1Br_2).
- The charge (without the sign) of one ion becomes the subscript of the other:

$$Ca^{2+} \quad Br^{1-} \qquad \text{gives} \qquad Ca_1Br_2 \qquad \text{or} \qquad CaBr_2$$

Reduce the subscripts to the smallest whole numbers that retain the ratio of ions. Thus, for example, from the ions Ca^{2+} and O^{2-} we have Ca_2O_2, which we reduce to the formula CaO (but see the footnote).

*Compounds of the mercury(I) ion, such as Hg_2Cl_2, and peroxides of the alkali metals, such as Na_2O_2, are the only two common exceptions. Their empirical formulas are HgCl and NaO, respectively.

SAMPLE PROBLEM 2.8 Determining Formulas of Binary Ionic Compounds

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PROBLEM Write empirical formulas for the compounds named in Sample Problem 2.7.
PLAN We write the empirical formula by finding the smallest number of each ion that gives the neutral compound. These numbers appear as *right subscripts* to the element symbol.
SOLUTION
(a) Mg^{2+} and N^{3-}; three Mg^{2+} ions (6+) balance two N^{3-} ions (6−): Mg_3N_2
(b) Cd^{2+} and I^-; one Cd^{2+} ion (2+) balances two I^- ions (2−): CdI_2
(c) Sr^{2+} and F^-; one Sr^{2+} ion (2+) balances two F^- ions (2−): SrF_2
(d) Cs^+ and S^{2-}; two Cs^+ ions (2+) balance one S^{2-} ion (2−): Cs_2S
COMMENT Note that ion charges do *not* appear in the compound formula. That is, for cadmium iodide, we do *not* write $Cd^{2+}I_2^-$.

FOLLOW-UP PROBLEM 2.8 Write the formulas of the compounds named in Follow-up Problem 2.7.

Compounds with Metals That Can Form More Than One Ion Many metals, particularly the transition elements (B groups), can form more than one ion, each with its own particular charge. Table 2.4 lists some examples, and the earlier Figure 2.18 shows their placement in the periodic table. Names of compounds containing these elements include a *Roman numeral within parentheses* immediately after the metal ion's name to indicate its ionic charge. For example, iron can form Fe^{2+} and Fe^{3+} ions. The two compounds that iron forms with chlorine are $FeCl_2$, named iron(II) chloride (spoken "iron two chloride"), and $FeCl_3$, named iron(III) chloride.

In common names, the Latin root of the metal is followed by either of two suffixes:

- The suffix *-ous* for the ion with the lower charge
- The suffix *-ic* for the ion with the higher charge

Thus, iron(II) chloride is also called ferr*ous* chloride and iron(III) chloride is fer*ric* chloride. (You can easily remember this naming relationship because there is an ***o*** in -***ous*** and ***lo****wer,* and an ***i*** in -***ic*** and *higher.*)

Table 2.4 Some Metals That Form More Than One Monatomic Ion*

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Element	Ion Formula	Systematic Name	Common (Trivial) Name
Chromium	Cr^{2+}	chromium(II)	chromous
	Cr^{3+}	**chromium(III)**	chromic
Cobalt	Co^{2+}	cobalt(II)	
	Co^{3+}	cobalt(III)	
Copper	**Cu^{+}**	**copper(I)**	cuprous
	Cu^{2+}	**copper(II)**	cupric
Iron	**Fe^{2+}**	**iron(II)**	ferrous
	Fe^{3+}	**iron(III)**	ferric
Lead	**Pb^{2+}**	**lead(II)**	
	Pb^{4+}	lead(IV)	
Mercury	Hg_2^{2+}	mercury(I)	mercurous
	Hg^{2+}	**mercury(II)**	mercuric
Tin	**Sn^{2+}**	**tin(II)**	stannous
	Sn^{4+}	tin(IV)	stannic

*Listed alphabetically by metal name; those in **boldface** are most common.

Table 2.5 Common Polyatomic Ions*

Formula	Name
Cations	
NH_4^+	**ammonium**
H_3O^+	**hydronium**
Anions	
CH_3COO^- (or **$C_2H_3O_2^-$**)	**acetate**
CN^-	cyanide
OH^-	**hydroxide**
ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	**chlorate**
ClO_4^-	**perchlorate**
NO_2^-	nitrite
NO_3^-	**nitrate**
MnO_4^-	**permanganate**
CO_3^{2-}	**carbonate**
HCO_3^-	**hydrogen carbonate** (or **bicarbonate**)
CrO_4^{2-}	chromate
$Cr_2O_7^{2-}$	**dichromate**
O_2^{2-}	peroxide
PO_4^{3-}	**phosphate**
HPO_4^{2-}	hydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate
SO_3^{2-}	sulfite
SO_4^{2-}	**sulfate**
HSO_4^-	hydrogen sulfate (or bisulfate)

*__Boldface__ ions are most common.

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SAMPLE PROBLEM 2.9 Determining Names and Formulas of Ionic Compounds of Elements That Form More Than One Ion

PROBLEM Give the systematic names for the formulas or the formulas for the names of the following compounds: **(a)** tin(II) fluoride; **(b)** CrI_3; **(c)** ferric oxide; **(d)** CoS.

SOLUTION **(a)** Tin(II) is Sn^{2+}; fluoride is F^-. Two F^- ions balance one Sn^{2+} ion: tin(II) fluoride is SnF_2. (The common name is stannous fluoride.)

(b) The anion is I^-, iodide, and the formula shows three I^-. Therefore, the cation must be Cr^{3+}, chromium(III): CrI_3 is chromium(III) iodide. (The common name is chromic iodide.)

(c) Ferric is the common name for iron(III), Fe^{3+}; oxide ion is O^{2-}. To balance the ionic charges, the formula of ferric oxide is Fe_2O_3. [The systematic name is iron(III) oxide.]

(d) The anion is sulfide, S^{2-}, which requires that the cation be Co^{2+}. The name is cobalt(II) sulfide.

FOLLOW-UP PROBLEM 2.9 Give the systematic names for the formulas or the formulas for the names of the following compounds: **(a)** lead(IV) oxide; **(b)** Cu_2S; **(c)** $FeBr_2$; **(d)** mercuric chloride.

Compounds Formed from Polyatomic Ions Ionic compounds in which one or both of the ions are polyatomic are very common. Table 2.5 gives the formulas and the names of some common polyatomic ions. Remember that *the polyatomic ion stays together as a charged unit*. The formula for potassium nitrate is KNO_3: each K^+ balances one NO_3^-. The formula for sodium carbonate is Na_2CO_3: two Na^+ balance one CO_3^{2-}. *When two or more of the same polyatomic ion are present in the formula unit, that ion appears in parentheses with the subscript written outside.* For example, calcium nitrate, which contains one Ca^{2+} and two NO_3^- ions, has the formula $Ca(NO_3)_2$. Parentheses and a subscript are *not* used unless *more than one* of the polyatomic ions is present; thus, sodium nitrate is $NaNO_3$, *not* $Na(NO_3)$.

Families of Oxoanions As Table 2.5 shows, most polyatomic ions are **oxoanions,** those in which an element, usually a nonmetal, is bonded to one or more oxygen atoms. There are several families of two or four oxoanions that differ only in the number of oxygen atoms. The following simple naming conventions are used with these ions.

With two oxoanions in the family:

- The ion with *more* O atoms takes the nonmetal root and the suffix *-ate*.
- The ion with *fewer* O atoms takes the nonmetal root and the suffix *-ite*.

For example, SO_4^{2-} is the sulf*ate* ion, and SO_3^{2-} is the sulf*ite* ion; similarly, NO_3^- is nitr*ate*, and NO_2^- is nitr*ite*.

With four oxoanions in the family (usually a halogen bonded to O), as Figure 2.19 shows:

- The ion with *most* O atoms has the prefix *per-*, the nonmetal root, and the suffix *-ate*.
- The ion with *one fewer* O atom has just the root and the suffix *-ate*.
- The ion with *two fewer* O atoms has just the root and the suffix *-ite*.
- The ion with *least (three fewer)* O atoms has the prefix *hypo-*, the root, and the suffix *-ite*.

For example, for the four chlorine oxoanions,

ClO_4^- is *per*chlor*ate*, ClO_3^- is chlor*ate*, ClO_2^- is chlor*ite*, ClO^- is *hypo*chlor*ite*

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No. of O atoms	Prefix	Root	Suffix
↑	per	*root*	ate
		root	ate
		root	ite
	hypo	*root*	ite

Figure 2.19 Naming oxoanions. Prefixes and suffixes indicate the number of O atoms in the anion.

Literature Marmorino, M. G. "Mnemonic for Oxoanion Formulas," *Chem. Educator* 10 **2005,** *6,* 415.

Hydrated Ionic Compounds Ionic compounds called **hydrates** have a specific number of water molecules associated with each formula unit. In their formulas, this number is shown after a centered dot. It is indicated in the systematic name by a Greek numerical prefix before the word *hydrate*. Table 2.6 shows these prefixes. For example, Epsom salt has the formula $MgSO_4{\cdot}7H_2O$ and the name magnesium sulfate *hepta*hydrate. Similarly, the mineral gypsum has the formula $CaSO_4{\cdot}2H_2O$ and the name calcium sulfate *di*hydrate. The water molecules, referred to as "waters of hydration," are part of the hydrate's structure. Heating can remove some or all of them, leading to a different substance. For example, when heated strongly, blue copper(II) sulfate pentahydrate ($CuSO_4{\cdot}5H_2O$) is converted to white copper(II) sulfate ($CuSO_4$).

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Table 2.6 Numerical Prefixes for Hydrates and Binary Covalent Compounds

Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

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SAMPLE PROBLEM 2.10 Determining Names and Formulas of Ionic Compounds Containing Polyatomic Ions

PROBLEM Give the systematic names for the formulas or the formulas for the names of the following compounds:
(a) $Fe(ClO_4)_2$ **(b)** Sodium sulfite **(c)** $Ba(OH)_2{\cdot}8H_2O$

SOLUTION **(a)** ClO_4^- is perchlorate, which has a 1− charge, so the cation must be Fe^{2+}. The name is iron(II) perchlorate. (The common name is ferrous perchlorate.)
(b) Sodium is Na^+; sulfite is SO_3^{2-}. Therefore, two Na^+ ions balance one SO_3^{2-} ion. The formula is Na_2SO_3. **(c)** Ba^{2+} is barium; OH^- is hydroxide. There are eight (*octa-*) water molecules in each formula unit. The name is barium hydroxide octahydrate.

FOLLOW-UP PROBLEM 2.10 Give the systematic names for the formulas or the formulas for the names of the following compounds:
(a) Cupric nitrate trihydrate **(b)** Zinc hydroxide **(c)** LiCN

SAMPLE PROBLEM 2.11 Recognizing Incorrect Names and Formulas of Ionic Compounds

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PROBLEM Something is wrong with the second part of each statement. Provide the correct name or formula.
(a) $Ba(C_2H_3O_2)_2$ is called barium diacetate.
(b) Sodium sulfide has the formula $(Na)_2SO_3$.
(c) Iron(II) sulfate has the formula $Fe_2(SO_4)_3$.
(d) Cesium carbonate has the formula $Cs_2(CO_3)$.

SOLUTION **(a)** The charge of the Ba^{2+} ion *must* be balanced by *two* $C_2H_3O_2^-$ ions, so the prefix *di-* is unnecessary. For ionic compounds, we do not indicate the number of ions with numerical prefixes. The correct name is barium acetate.
(b) Two mistakes occur here. The sodium ion is monatomic, so it does *not* require parentheses. The sulfide ion is S^{2-}, *not* SO_3^{2-} (called "sulfite"). The correct formula is Na_2S.
(c) The Roman numeral refers to the charge of the ion, *not* the number of ions in the formula. Fe^{2+} is the cation, so it requires one SO_4^{2-} to balance its charge. The correct formula is $FeSO_4$.
(d) Parentheses are *not* required when only one polyatomic ion of a kind is present. The correct formula is Cs_2CO_3.

FOLLOW-UP PROBLEM 2.11 State why the second part of each statement is incorrect, and correct it:
(a) Ammonium phosphate is $(NH_3)_4PO_4$.
(b) Aluminum hydroxide is $AlOH_3$.
(c) $Mg(HCO_3)_2$ is manganese(II) carbonate.
(d) $Cr(NO_3)_3$ is chromic(III) nitride.
(e) $Ca(NO_2)_2$ is cadmium nitrate.

Acid Names from Anion Names Acids are an important group of hydrogen-containing compounds that have been used in chemical reactions since before alchemical times. In the laboratory, acids are typically used in water solution. When naming them and writing their formulas, we consider them as anions connected to the number of hydrogen ions (H^+) needed for charge neutrality. The two common types of acids are binary acids and oxoacids:

1. *Binary acid* solutions form when certain gaseous compounds dissolve in water. For example, when gaseous hydrogen chloride (HCl) dissolves in water, it forms a solution whose name consists of the following parts:

   Prefix *hydro-* + nonmetal *root* + suffix *-ic* + separate word *acid*
   hydro + chlor + ic + acid

   or *hydrochloric acid*. This naming pattern holds for many compounds in which hydrogen combines with an anion that has an *-ide* suffix.
2. *Oxoacid* names are similar to those of the oxoanions, except for two suffix changes:
   - *-ate* in the anion becomes *-ic* in the acid
   - *-ite* in the anion becomes *-ous* in the acid

   The oxoanion prefixes *hypo-* and *per-* are kept. Thus,

   BrO_4^- is *per*brom*ate*, and $HBrO_4$ is *per*brom*ic* acid
   IO_2^- is iod*ite*, and HIO_2 is iod*ous* acid

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SAMPLE PROBLEM 2.12 Determining Names and Formulas of Anions and Acids

PROBLEM Name the following anions and give the names and formulas of the acids derived from them: **(a)** Br^-; **(b)** IO_3^-; **(c)** CN^-; **(d)** SO_4^{2-}; **(e)** NO_2^-.

SOLUTION **(a)** The anion is bromide; the acid is hydrobromic acid, HBr.

(b) The anion is iodate; the acid is iodic acid, HIO_3.

(c) The anion is cyanide; the acid is hydrocyanic acid, HCN.

(d) The anion is sulfate; the acid is sulfuric acid, H_2SO_4. (In this case, the suffix is added to the element name *sulfur,* not to the root, *sulf-*.)

(e) The anion is nitrite; the acid is nitrous acid, HNO_2.

COMMENT We added *two* H^+ ions to the sulfate ion to obtain sulfuric acid because SO_4^{2-} has a 2− charge.

FOLLOW-UP PROBLEM 2.12 Write the formula for the name or name for the formula of each acid: **(a)** chloric acid; **(b)** HF; **(c)** acetic acid; **(d)** sulfurous acid; **(e)** HBrO.

Names and Formulas of Binary Covalent Compounds

Binary covalent compounds are formed by the combination of two elements, usually nonmetals. Some are so familiar, such as ammonia (NH_3), methane (CH_4), and water (H_2O), we use their common names, but most are named in a systematic way:

1. The element with the lower group number in the periodic table is the first word in the name; the element with the higher group number is the second word. (*Exception:* When the compound contains oxygen and any of the halogens chlorine, bromine, and iodine, the halogen is named first.)
2. If both elements are in the same group, the one with the higher period number is named first.
3. The second element is named with its root and the suffix *-ide*.
4. Covalent compounds have Greek numerical prefixes (see Table 2.6) to indicate the number of atoms of each element in the compound. The first word has a prefix *only* when more than one atom of the element is present; the second word *usually* has a numerical prefix.

SAMPLE PROBLEM 2.13 Determining Names and Formulas of Binary Covalent Compounds

PROBLEM **(a)** What is the formula of carbon disulfide?
(b) What is the name of PCl_5?
(c) Give the name and formula of the compound whose molecules each consist of two N atoms and four O atoms.

SOLUTION **(a)** The prefix *di-* means "two." The formula is CS_2.
(b) P is the symbol for phosphorus; there are five chlorine atoms, which is indicated by the prefix *penta-*. The name is phosphorus pentachloride.
(c) Nitrogen (N) comes first in the name (lower group number). The compound is dinitrogen tetraoxide, N_2O_4.

FOLLOW-UP PROBLEM 2.13 Give the name or formula for **(a)** SO_3; **(b)** SiO_2; **(c)** dinitrogen monoxide; **(d)** selenium hexafluoride.

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SAMPLE PROBLEM 2.14 Recognizing Incorrect Names and Formulas of Binary Covalent Compounds

PROBLEM Explain what is wrong with the name or formula in the second part of each statement and correct it: **(a)** SF_4 is monosulfur pentafluoride. **(b)** Dichlorine heptaoxide is Cl_2O_6. **(c)** N_2O_3 is dinitrotrioxide.

SOLUTION **(a)** There are two mistakes. *Mono-* is not needed if there is only one atom of the first element, and the prefix for four is *tetra-*, not *penta-*. The correct name is sulfur tetrafluoride.
(b) The prefix *hepta-* indicates seven, not six. The correct formula is Cl_2O_7.
(c) The full name of the first element is needed, and a space separates the two element names. The correct name is dinitrogen trioxide.

FOLLOW-UP PROBLEM 2.14 Explain what is wrong with the second part of each statement and correct it: **(a)** S_2Cl_2 is disulfurous dichloride. **(b)** Nitrogen monoxide is N_2O. **(c)** $BrCl_3$ is trichlorine bromide.

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An Introduction to Naming Organic Compounds

Organic compounds typically have complex structural formulas that consist of chains, branches, and/or rings of carbon atoms bonded to hydrogen atoms and, often, to atoms of oxygen, nitrogen, and a few other elements. At this point, we'll look at one or two basic principles for naming them. Much more on the rules of organic nomenclature appears in Chapter 15.

Hydrocarbons, the simplest type of organic compound, contain *only* carbon and hydrogen. *Alkanes* are the simplest type of hydrocarbon; many function as important fuels, such as methane, propane, butane, and the mixture of alkanes in gasoline. The simplest alkanes to name are the *straight-chain alkanes* because the carbon chains have no branches. Alkanes are named with a *root,* based on the number of C atoms in the chain, followed by the suffix *-ane*. Table 2.7 gives the names, molecular formulas, and space-filling models (discussed shortly) of the first 10 straight-chain alkanes. Note that the roots of the four smallest ones are new, but those for the larger ones are the same as the Greek prefixes shown in Table 2.6.

Alkanes (and other organic compounds) *with* branches have a *prefix* in the name as well. *The prefix names the length of the branch and numbers the carbon atom in the main chain that the branch is attached to.* A prefix consists of a root plus the ending *-yl*. Thus, for example, the compound with a one-carbon ("meth") branch attached to the second carbon of the main chain of butane is *2-methylbutane,* where "2-methyl" is the prefix (see margin).

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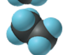
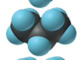
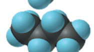
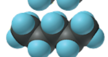
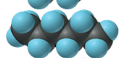
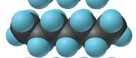
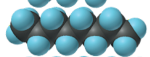
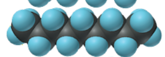
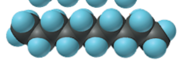
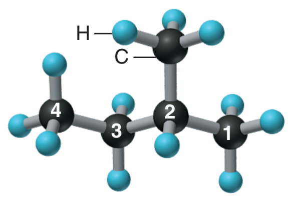
Table 2.7 The First 10 Straight-Chain Alkanes

Name (Formula)	Model
Methane (CH_4)	
Ethane (C_2H_6)	
Propane (C_3H_8)	
Butane (C_4H_{10})	
Pentane (C_5H_{12})	
Hexane (C_6H_{14})	
Heptane (C_7H_{16})	
Octane (C_8H_{18})	
Nonane (C_9H_{20})	
Decane ($C_{10}H_{22}$)	



Ball-and-stick model of 2-methylbutane

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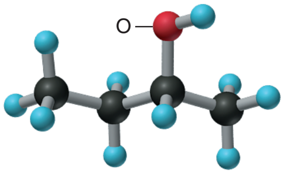
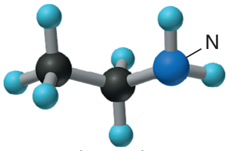
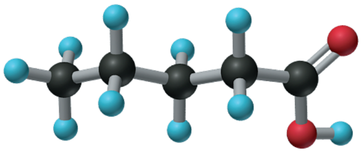


2-butanol



ethanamine


pentanoic acid

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Organic chemistry application

Organic compounds other than alkanes have names derived from a particular region of the molecule, called the *functional group,* which consists of one or a few atoms bonded in a specific way. *The functional group determines how the compound reacts.* The *alcohol* functional group is a hydroxyl group, O—H, in place of one of the H atoms in the hydrocarbon; an *amine* has an amino group, $—NH_2$; a *carboxylic acid* has a carboxyl group, —COOH; and so forth. Each functional group has its own suffix. Thus, the compound with a hydroxyl group attached to the second carbon in butane is called *2-butanol* (see margin); the compound with an amino group bonded to a two-carbon chain is *ethanamine;* the compound with a carboxyl group bonded to a four-carbon chain is *pentanoic acid* (the C of the carboxyl group counts as one of the carbons). We'll examine how the different functional groups react in Chapter 15.

Molecular Masses from Chemical Formulas

In Section 2.5, we calculated the atomic mass of an element. Using the periodic table and the formula of a compound to see the number of atoms of each element, we calculate the **molecular mass** (also called *molecular weight*) of a formula unit of the compound as the sum of the atomic masses:

$$\text{Molecular mass} = \text{sum of atomic masses} \tag{2.3}$$

The molecular mass of a water molecule (using atomic masses to four significant figures from the periodic table) is

$$\begin{aligned}\text{Molecular mass of } H_2O &= (2 \times \text{atomic mass of H}) + (1 \times \text{atomic mass of O})\\ &= (2 \times 1.008\text{ amu}) + 16.00\text{ amu} = 18.02\text{ amu}\end{aligned}$$

Ionic compounds are treated the same, but because they do not consist of molecules, we use the term **formula mass** for an ionic compound. To calculate its formula mass, *the number of atoms of each element inside the parentheses is multiplied by the subscript outside the parentheses.* For barium nitrate, $Ba(NO_3)_2$,

$$\begin{aligned}&\text{Formula mass of } Ba(NO_3)_2\\ &\quad = (1 \times \text{atomic mass of Ba}) + (2 \times \text{atomic mass of N}) + (6 \times \text{atomic mass of O})\\ &\quad = 137.3\text{ amu} + (2 \times 14.01\text{ amu}) + (6 \times 16.00\text{ amu}) = 261.3\text{ amu}\end{aligned}$$

Atomic, not ionic, masses are used because electron loss equals electron gain in the compound, so electron mass is balanced. In the next two problems, the name or molecular depiction is used to find a compound's formula and molecular mass.

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SAMPLE PROBLEM 2.15 Calculating the Molecular Mass of a Compound

PROBLEM Using data in the periodic table, calculate the molecular (or formula) mass of:
(a) Tetraphosphorus trisulfide **(b)** Ammonium nitrate

PLAN We first write the formula, then multiply the number of atoms (or ions) of each element by its atomic mass, and find the sum.

SOLUTION **(a)** The formula is P_4S_3.

$$\begin{aligned}\text{Molecular mass} &= (4 \times \text{atomic mass of P}) + (3 \times \text{atomic mass of S})\\ &= (4 \times 30.97\text{ amu}) + (3 \times 32.07\text{ amu}) = \boxed{220.09\text{ amu}}\end{aligned}$$

(b) The formula is NH_4NO_3. We count the total number of N atoms even though they belong to different ions:

$$\begin{aligned}&\text{Formula mass}\\ &\quad = (2 \times \text{atomic mass of N}) + (4 \times \text{atomic mass of H}) + (3 \times \text{atomic mass of O})\\ &\quad = (2 \times 14.01\text{ amu}) + (4 \times 1.008\text{ amu}) + (3 \times 16.00\text{ amu}) = \boxed{80.05\text{ amu}}\end{aligned}$$

CHECK You can often find large errors by rounding atomic masses to the nearest 5 and adding: **(a)** $(4 \times 30) + (3 \times 30) = 210 \approx 220.09$. The sum has two decimal places because the atomic masses have two. **(b)** $(2 \times 15) + 4 + (3 \times 15) = 79 \approx 80.05$.

FOLLOW-UP PROBLEM 2.15 What is the formula and molecular (or formula) mass of each of the following compounds: **(a)** hydrogen peroxide; **(b)** cesium chloride; **(c)** sulfuric acid; **(d)** potassium sulfate?

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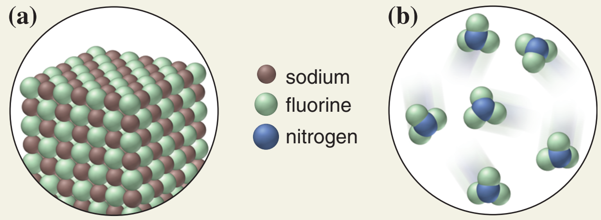
SAMPLE PROBLEM 2.16 Using Molecular Depictions to Determine Formula, Name, and Mass

PROBLEM Each circle contains a representation of a binary compound. Determine its formula, name, and molecular (formula) mass.



PLAN Each of the compounds contains only two elements, so to find the formula, we find the simplest whole-number ratio of one atom to the other. Then we determine the name and formula (see Sample Problems 2.7–2.9 and 2.13) and the mass (see Sample Problem 2.15).

SOLUTION **(a)** There is one brown (sodium) for each green (fluorine), so the formula is NaF. A metal and nonmetal form an ionic compound, in which the metal is named first: sodium fluoride.

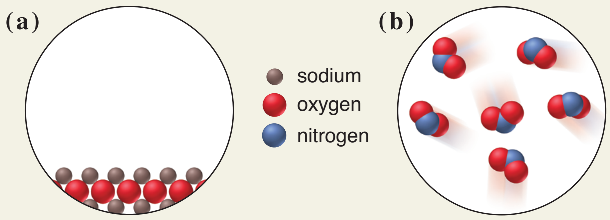
$$\begin{aligned}\text{Formula mass} &= (1 \times \text{atomic mass of Na}) + (1 \times \text{atomic mass of F})\\ &= 22.99\ \text{amu} + 19.00\ \text{amu} = 41.99\ \text{amu}\end{aligned}$$

(b) There are three green (fluorine) for each blue (nitrogen), so the formula is NF_3. Two nonmetals form a covalent compound. Nitrogen has a lower group number, so it is named first: nitrogen trifluoride.

$$\begin{aligned}\text{Molecular mass} &= (1 \times \text{atomic mass of N}) + (3 \times \text{atomic mass of F})\\ &= 14.01\ \text{amu} + (3 \times 19.00\ \text{amu}) = 71.01\ \text{amu}\end{aligned}$$

CHECK **(a)** For binary ionic compounds, we predict ionic charges from the periodic table (see Figure 2.13). Na forms a 1+ ion, and F forms a 1− ion, so the charges balance with one Na^+ per F^-. Also, ionic compounds are solids, consistent with the picture. **(b)** Covalent compounds often occur as individual molecules, as in the picture. Rounding in (a) gives 25 + 20 = 45; in (b), we get 15 + (3 × 20) = 75, so there are no large errors.

FOLLOW-UP PROBLEM 2.16 Each circle contains a representation of a binary compound. Determine its name, formula, and molecular (formula) mass.



The Gallery on the next page shows some of the ways that chemists picture molecules and the enormous range of molecular sizes.

Section Summary

Chemical formulas describe the simplest atom ratio (empirical formula), actual atom number (molecular formula), and atom arrangement (structural formula) of one unit of a compound. • An ionic compound is named with cation first and anion second. For metals that can form more than one ion, the charge is shown with a Roman numeral. • Oxoanions have suffixes, and sometimes prefixes, attached to the root of the element name to indicate the number of oxygen atoms. • Names of hydrates give the number of associated water molecules with a numerical prefix. • Acid names are based on anion names. • Covalent compounds have as the first word of the name the element that is farther left or lower down in the periodic table, and prefixes show the numbers of each atom. • The molecular (or formula) mass of a compound is the sum of the atomic masses in the formula. • Molecules are three-dimensional objects that range in size from H_2 to biological and synthetic macromolecules.

Picturing Molecules

The most exciting thing about learning chemistry is training your mind to imagine a molecular world, one filled with tiny objects of various shapes. Molecules are depicted in a variety of useful ways, as shown at right for the water molecule:

Chemical formulas show only the relative numbers of atoms.

Electron-dot and *bond-line formulas* show a bond between atoms as either a pair of dots or a line.

Ball-and-stick models show atoms as spheres and bonds as sticks, with accurate angles and relative sizes, but distances are exaggerated.

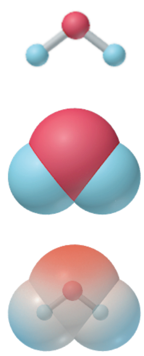
Space-filling models are accurately scaled-up versions of molecules, but they do not show bonds.

Electron-density models show the ball-and-stick model within the space-filling shape and color the regions of high *(red)* and low *(blue)* electron charge.

H_2O

H:O:H

H–O–H



All molecules are minute, with their relative sizes depending on composition. A water molecule is small because it consists of only three atoms. Many air pollutants, such as ozone, carbon monoxide, and nitrogen dioxide, also consist of small molecules.

Ozone (O_3, 48.00 amu) contributes to smog; natural component of stratosphere that absorbs harmful solar radiation.

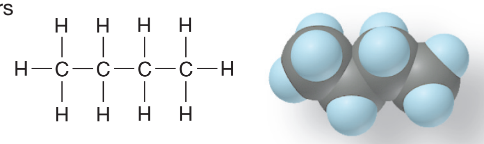
Carbon monoxide (CO, 28.01 amu), toxic component of car exhaust and cigarette smoke.

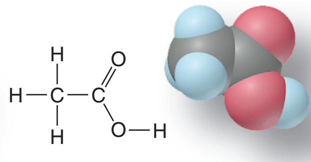
C≡O

Nitrogen dioxide (NO_2, 46.01 amu) forms from nitrogen monoxide and contributes to smog and acid rain.

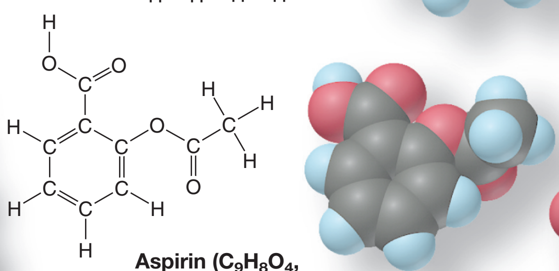
Many household chemicals, such as butane, acetic acid, and aspirin, consist of somewhat larger molecules. The biologically essential molecule heme is larger still.

Butane (C_4H_{10}, 58.12 amu), fuel for cigarette lighters and camping stoves.

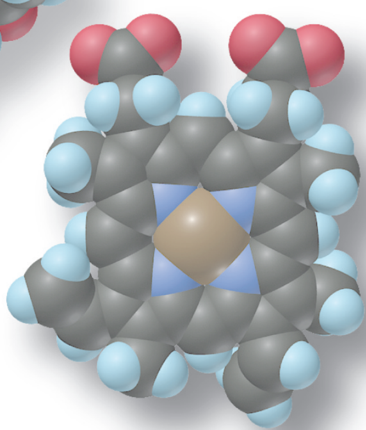




Acetic acid (CH_3COOH, 60.05 amu), component of vinegar.

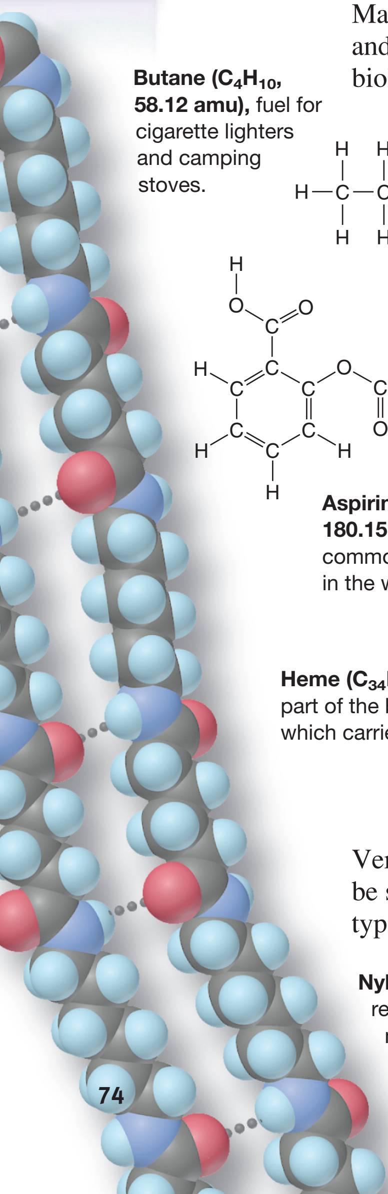


Aspirin ($C_9H_8O_4$, 180.15 amu), most common pain reliever in the world.

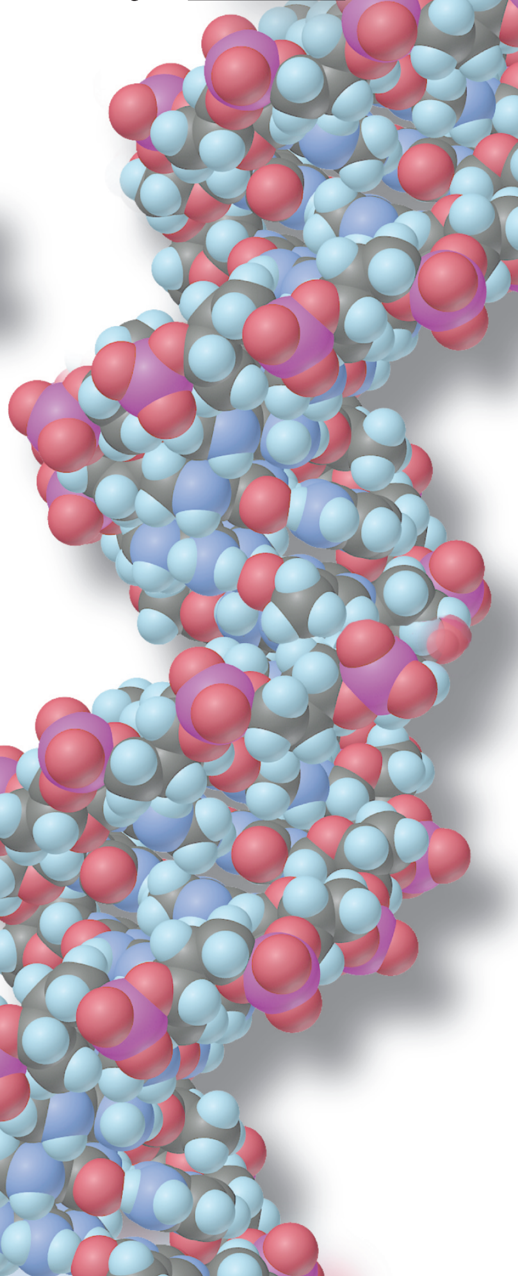


Heme ($C_{34}H_{32}FeN_4O_4$, 616.49 amu), part of the blood protein hemoglobin, which carries oxygen through the body.

Very large molecules, called macromolecules, can be synthetic, like nylon, or natural, like DNA, and typically consist of thousands of atoms.



Nylon-66 (~15,000 amu), relatively small, synthetic macromolecule used to make textiles and tires.



Deoxyribonucleic acid (DNA, ~10,000,000 amu), cellular macromolecule that contains genetic information.

2.9 MIXTURES: CLASSIFICATION AND SEPARATION

Although chemists pay a great deal of attention to pure substances, this form of matter almost never occurs around us. In the natural world, *matter usually occurs as mixtures*. A sample of clean air, for example, consists of many elements and compounds physically mixed together, including oxygen (O_2), nitrogen (N_2), carbon dioxide (CO_2), the six noble gases [Group 8A(18)], and water vapor (H_2O). The oceans are complex mixtures of dissolved ions and covalent substances, including Na^+, Mg^{2+}, Cl^-, SO_4^{2-}, O_2, CO_2, and of course H_2O. Rocks and soils are mixtures of numerous compounds—such as calcium carbonate ($CaCO_3$), silicon dioxide (SiO_2), aluminum oxide (Al_2O_3), and iron(III) oxide (Fe_2O_3)—perhaps a few elements (gold, silver, and carbon in the form of diamond), and petroleum and coal, which are complex mixtures themselves. Living things contain thousands of substances: carbohydrates, lipids, proteins, nucleic acids, and many simpler ionic and covalent compounds.

There are two broad classes of mixtures. A **heterogeneous mixture** has one or more visible boundaries between the components. Thus, its composition is *not* uniform. Many rocks are heterogeneous, showing individual grains and flecks of different minerals. In some cases, as in milk and blood, the boundaries can be seen only with a microscope. A **homogeneous mixture** has no visible boundaries because the components are mixed as individual atoms, ions, and molecules. Thus, its composition *is* uniform, unvarying from one region to another. A mixture of sugar dissolved in water is homogeneous, for example, because the sugar molecules and water molecules are uniformly intermingled on the molecular level. We have no way to tell visually whether an object is a substance (element or compound) or a homogeneous mixture.

A homogeneous mixture is also called a **solution.** Although we usually think of solutions as liquid, they can exist in all three physical states. For example, air is a gaseous solution of mostly oxygen and nitrogen molecules, and wax is a solid solution of several fatty substances. Solutions in water, called **aqueous solutions,** are especially important in chemistry and comprise a major portion of the environment and of all organisms.

Recall that mixtures differ fundamentally from compounds in three ways: (1) the proportions of the components can vary; (2) the individual properties of the components are observable; and (3) the components can be separated by physical means. In some cases, if we apply enough energy to the components of the mixture, they react with each other chemically and form a compound, after which their individual properties are no longer observable. Figure 2.20 shows such a case with a mixture of iron and sulfur.

In order to investigate the properties of substances, chemists have devised many procedures for separating a mixture into its component elements and compounds. Indeed, the laws and models of chemistry could never have been formulated without this ability. Many of Dalton's critics, who thought they had found compounds with varying composition, were unknowingly studying mixtures! The Tools of the Laboratory essay on the next two pages describes some of the more common laboratory separation methods.

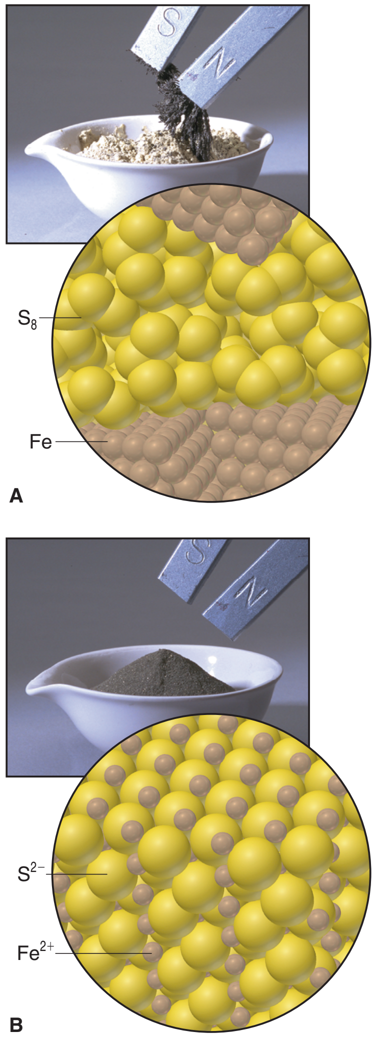


Figure 2.20 The distinction between mixtures and compounds. A, A *mixture* of iron and sulfur can be separated with a magnet because only the iron is magnetic. The blow-up shows separate regions of the two elements. **B,** After strong heating, the *compound* iron(II) sulfide forms, which is no longer magnetic. The blow-up shows the structure of the compound, in which there are no separate regions of the elements.

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Section Summary

Heterogeneous mixtures have visible boundaries between the components. • Homogeneous mixtures have no visible boundaries because mixing occurs at the molecular level. A solution is a homogeneous mixture and can occur in any physical state. • Components of mixtures (unlike those of compounds) can have variable proportions, can be separated physically, and retain their properties. • Common physical separation processes include filtration, crystallization, extraction, chromatography, and distillation.

Tools of the Laboratory

Basic Separation Techniques

Some of the most challenging and time-consuming laboratory procedures involve separating mixtures and purifying the components. Several common separation techniques are described here. All of these methods depend on the *physical properties* of the substances in the mixture; no chemical changes occur.

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Filtration separates the components of a mixture on the basis of *differences in particle size*. It is used most often to separate a liquid (smaller particles) from a solid (larger particles). Figure B2.3 shows simple filtration of a solid reaction product. In vacuum filtration, reduced pressure within the flask speeds the flow of the liquid through the filter. Filtration is a key step in the purification of the tap water you drink.

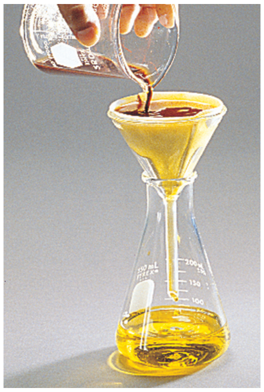

Figure B2.3 Filtration.

Crystallization is based on *differences in solubility*. The *solubility* of a substance is the amount that dissolves in a fixed volume of solvent at a given temperature. The result shown in Figure B2.4 applies the fact that many substances are more soluble in hot solvent than in cold. The purified compound crystallized as the solution was cooled. Key substances in computer chips and other electronic devices are purified by a type of crystallization.

Figure B2.4 Crystallization.

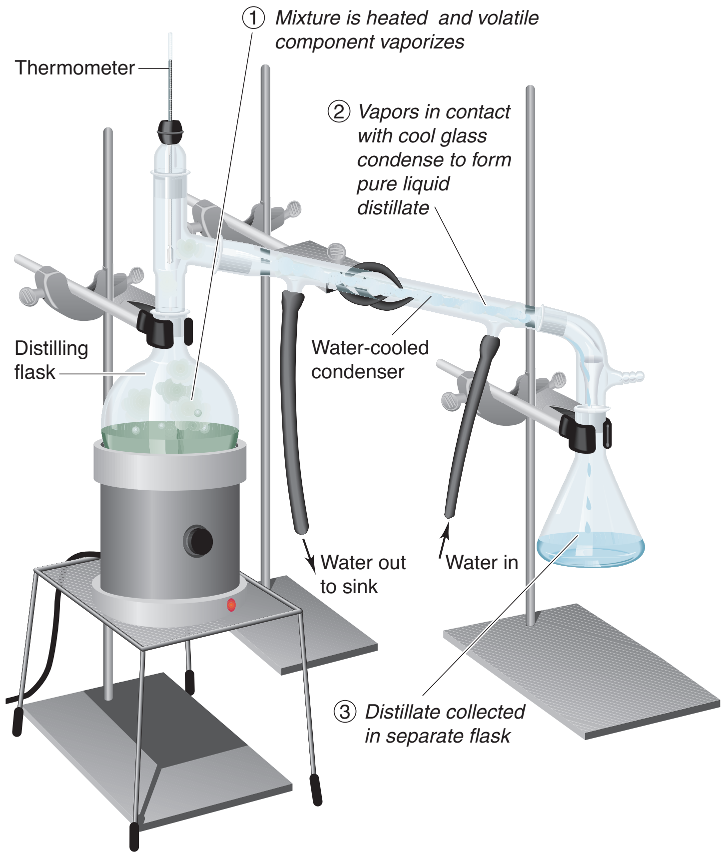

Figure B2.5 Distillation. ARIS Presentation Center, Art Library

Distillation separates components through *differences in **volatility***, the tendency of a substance to become a gas. Ether, for example, is more volatile than water, which is much more volatile than sodium chloride. The simple distillation apparatus shown in Figure B2.5 is used to separate components with *large* differences in volatility, such as water from dissolved ionic compounds. As the mixture boils, the vapor is richer in the more volatile component, which is condensed and collected separately. Separating components with small volatility differences requires many vaporization-condensation steps (discussed in Chapter 13).

Extraction is also based on *differences in solubility.* In a typical procedure, a natural (often plant or animal) material is ground in a blender with a solvent that extracts (dissolves) soluble compound(s) embedded in insoluble material. This extract is separated further by the addition of a second solvent that does not dissolve in the first. After shaking in a separatory funnel, some components are extracted into the new solvent. Figure B2.6 shows the extraction of plant pigments from water into hexane, an organic solvent.

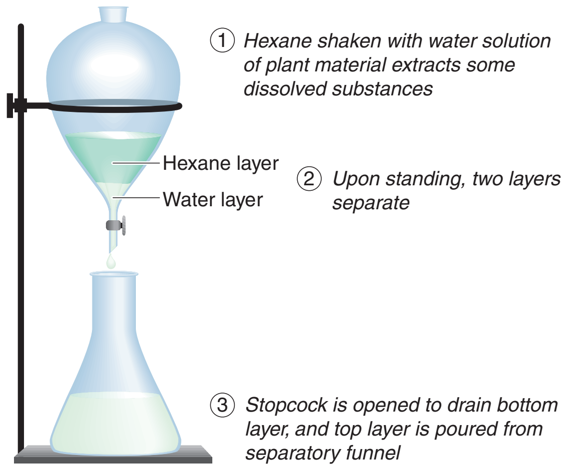

Figure B2.6 Extraction. ARIS Presentation Center, Art Library

Chromatography is a third technique based on *differences in solubility.* The mixture is dissolved in a gas or liquid called the *mobile phase,* and the components are separated as this phase moves over a solid (or viscous liquid) surface called the *stationary phase*. A component with low solubility in the stationary phase spends less time there, thus moving faster than a component that is highly soluble in that phase. Figure B2.7 depicts the separation of a mixture of pigments in ink.

Many types of chromatography are used to separate a wide variety of substances, from simple gases to biological macromolecules. In *gas-liquid chromatography (GLC),* the mobile phase is an inert gas, such as helium, that carries the previously vaporized components into a long tube that contains the stationary phase (Figure B2.8, part A). The components emerge separately and reach a detector to create a chromatogram. A typical chromatogram has numerous peaks of specific position and height, each of which represents the amount of a given component (Figure B2.8, part B).

The principle of *high-performance (high-pressure) liquid chromatography (HPLC)* is very similar. However, in this technique the mixture is not vaporized, so a more diverse group of components, which may include nonvolatile compounds, can be separated (Figure B2.9).

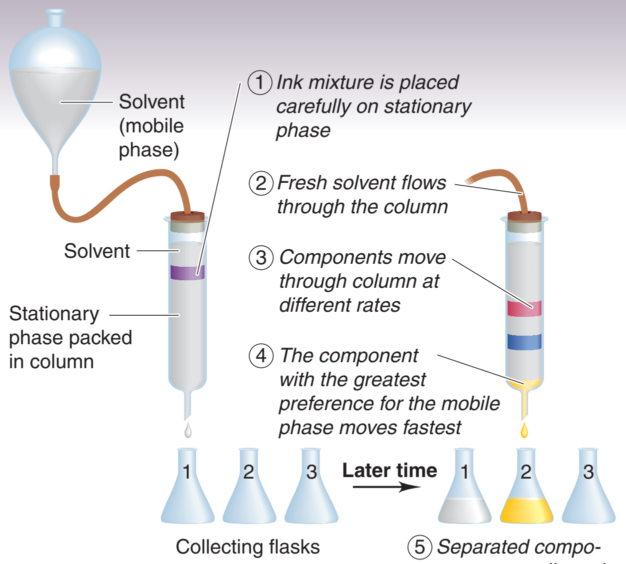

Figure B2.7 Procedure for column chromatography.

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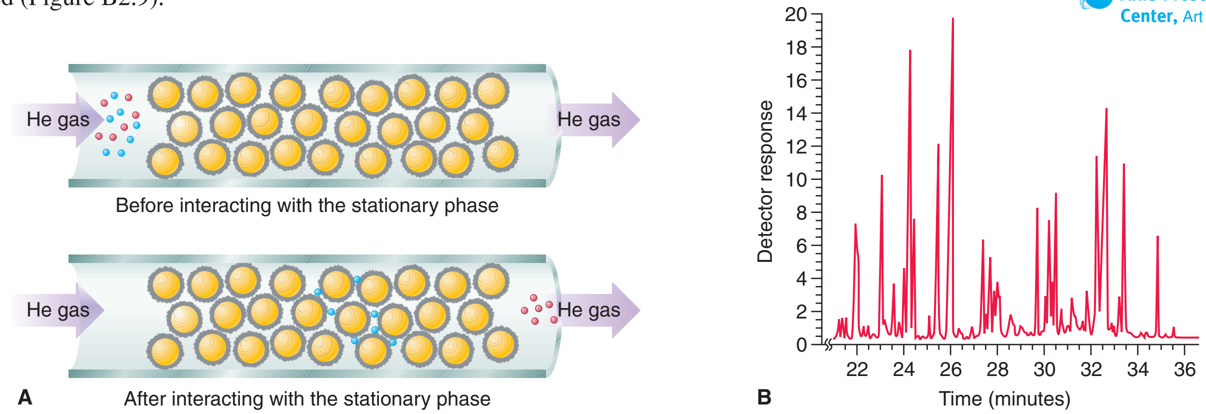

Figure B2.8 Principle of gas-liquid chromatography (GLC). A, The mobile phase *(purple arrow)* carries the sample mixture into a tube packed with the stationary phase *(gray outline on yellow spheres),* and each component dissolves in the stationary phase to a different extent. A component *(red)* that dissolves less readily than another *(blue)* emerges from the tube sooner. **B,** A typical gas-liquid chromatogram of a complex mixture displays each component as a peak.

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Figure B2.9 A high-performance liquid chromatograph.

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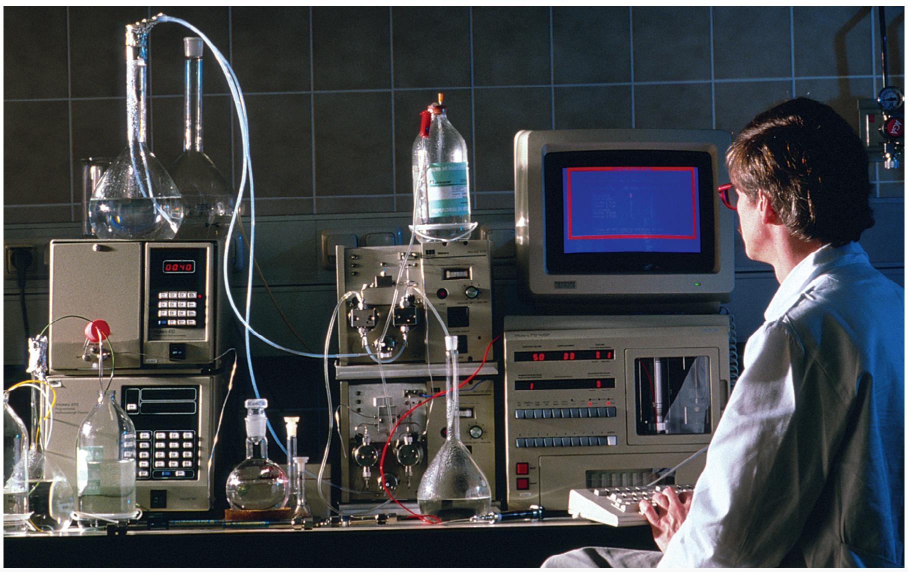

Chapter Perspective

An understanding of matter at the observable and atomic levels is the essence of chemistry. In this chapter, you have learned how matter is classified in terms of its composition and how it is named in words and formulas, which are major steps toward that understanding. Figure 2.21 provides a visual review of many key terms and ideas in this chapter. In Chapter 3, we explore one of the central quantitative ideas in chemistry: how the observable amount of a substance relates to the number of atoms, molecules, or ions that make it up.

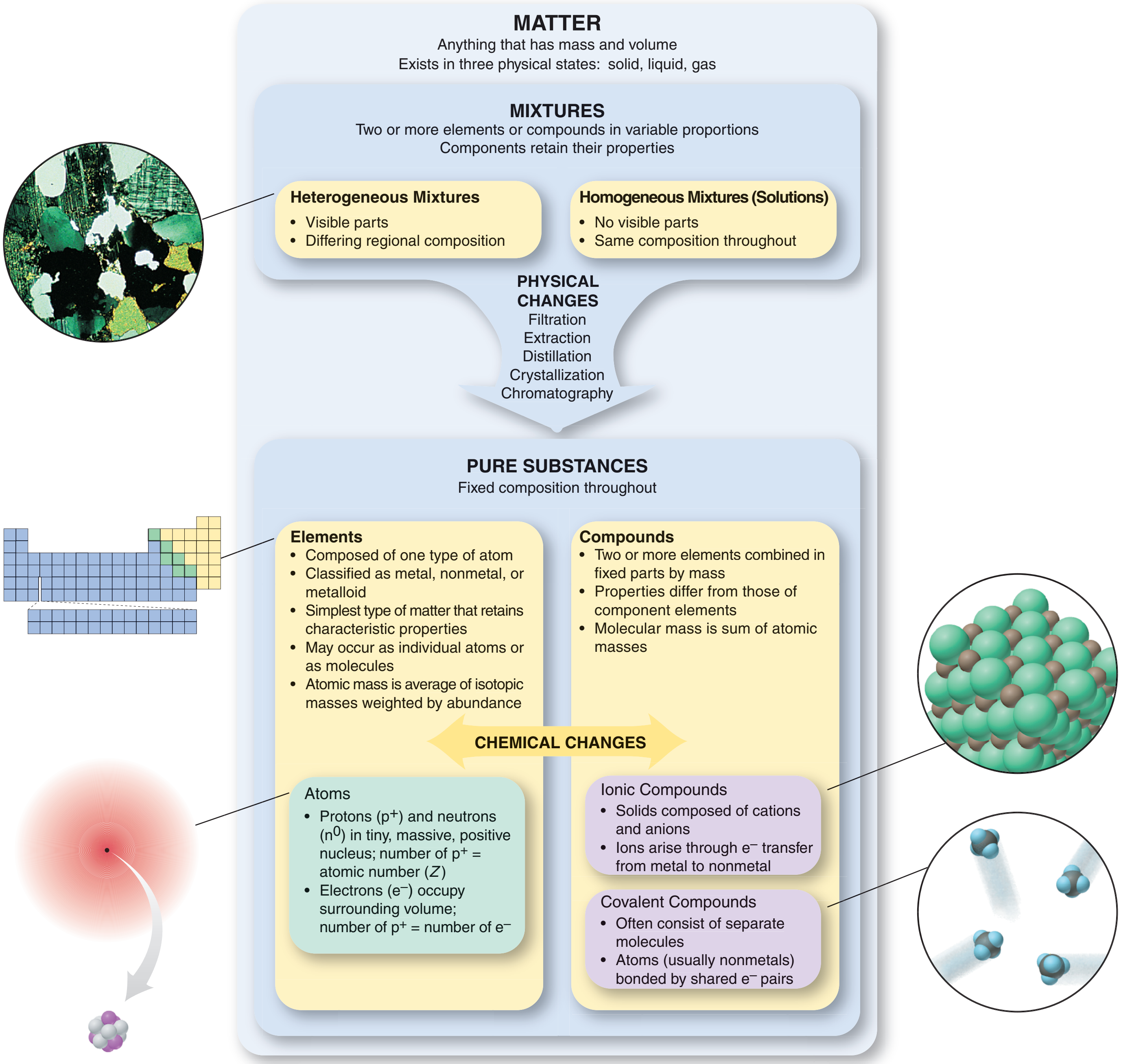

Figure 2.21 The classification of matter from a chemical point of view. Mixtures are separated by physical changes into elements and compounds. Chemical changes are required to convert elements into compounds, and vice versa.

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CHAPTER REVIEW GUIDE

The following sections provide many aids to help you study this chapter. (Numbers in parentheses refer to pages, unless noted otherwise.)

Learning Objectives

These are concepts and skills you should know after studying this chapter.

Relevant section and/or sample problem (SP) numbers appear in parentheses.

Understand These Concepts

1. The defining characteristics of the three types of matter—element, compound, and mixture—on the macroscopic and atomic levels (2.1)
2. The significance of the three mass laws—mass conservation, definite composition, and multiple proportions (2.2)
3. The postulates of Dalton's atomic theory and how it explains the mass laws (2.3)
4. The major contribution of experiments by Thomson, Millikan, and Rutherford concerning atomic structure (2.4)
5. The structure of the atom, the main features of the subatomic particles, and the importance of isotopes (2.5)
6. The format of the periodic table and general location and characteristics of metals, metalloids, and nonmetals (2.6)
7. The essential features of ionic and covalent bonding and the distinction between them (2.7)
8. The types of mixtures and their properties (2.9)

Master These Skills

1. Distinguishing elements, compounds, and mixtures at the atomic scale (SP 2.1)
2. Using the mass ratio of element to compound to find the mass of an element in a compound (SP 2.2)
3. Visualizing the mass laws (SP 2.3)
4. Using atomic notation to express the subatomic makeup of an isotope (SP 2.4)
5. Calculating an atomic mass from isotopic composition (SP 2.5)
6. Predicting the monatomic ion formed from a main-group element (SP 2.6)
7. Naming and writing the formula of an ionic compound formed from the ions in Tables 2.3 to 2.5 (SP 2.7–2.12, 2.16)
8. Naming and writing the formula of an inorganic binary covalent compound (SP 2.13, 2.14, 2.16)
9. Calculating the molecular or formula mass of a compound (SP 2.15)

Key Terms

These important terms appear in boldface in the chapter and are defined again in the Glossary.

Section 2.1
element (41)
substance (41)
molecule (42)
compound (42)
mixture (42)

Section 2.2
law of mass conservation (44)
law of definite (or constant) composition (44)
fraction by mass (mass fraction) (44)
percent by mass (mass percent, mass %) (44)
law of multiple proportions (46)

Section 2.3
atom (47)

Section 2.4
cathode ray (49)
nucleus (51)

Section 2.5
proton (p^+) (52)
neutron (n^0) (52)
electron (e^-) (52)
atomic number (Z) (53)
mass number (A) (53)
atomic symbol (53)
isotope (53)
atomic mass unit (amu) (54)
dalton (Da) (54)
mass spectrometry (54)
isotopic mass (54)
atomic mass (54)

Section 2.6
periodic table of the elements (57)
period (57)
group (57)
metal (58)
nonmetal (58)
metalloid (semimetal) (58)

Section 2.7
ionic compound (60)
covalent compound (60)
chemical bond (60)
ion (60)
binary ionic compound (60)
cation (60)
anion (60)
monatomic ion (60)
covalent bond (62)
polyatomic ion (63)

Section 2.8
chemical formula (64)
empirical formula (64)
molecular formula (65)
structural formula (65)
formula unit (66)
oxoanion (68)
hydrate (69)
binary covalent compound (70)
molecular mass (72)
formula mass (72)

Section 2.9
heterogeneous mixture (75)
homogeneous mixture (75)
solution (75)
aqueous solution (75)
filtration (76)
crystallization (76)
distillation (76)
volatility (76)
extraction (76)
chromatography (77)

Key Equations and Relationships

Numbered and screened concepts are listed for you to refer to or memorize.

2.1 Finding the mass of an element in a given mass of compound (45):

Mass of element in sample

$$= \text{mass of compound in sample} \times \frac{\text{mass of element}}{\text{mass of compound}}$$

2.2 Calculating the number of neutrons in an atom (53):

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}$$

or

$$N = A - Z$$

2.3 Determining the molecular mass of a formula unit of a compound (72):

$$\text{Molecular mass} = \text{sum of atomic masses}$$

Highlighted Figures and Tables

These figures (F) and tables (T) provide a visual review of key ideas.

Entries in **bold** contain frequently used data.

F2.1 Elements, compounds, and mixtures on atomic scale (42)
F2.7 General features of the atom (52)
T2.2 Properties of the three key subatomic particles (53)
F2.9 The modern periodic table (58)
F2.12 Factors that influence the strength of ionic bonding (61)
F2.13 The relationship of ions formed to the nearest noble gas (61)
F2.14 Formation of a covalent bond between two H atoms (62)
F2.18 Some common monatomic ions of the elements (65)
T2.3 Common monatomic ions (65)
T2.4 Some metals that form more than one monatomic ion (67)
T2.5 Common polyatomic ions (68)
T2.6 Numerical prefixes for hydrates and binary covalent compounds (69)
F2.21 Classification of matter from a chemical point of view (78)

Brief Solutions to FOLLOW-UP PROBLEMS

Compare your solutions to these calculation steps and answers.

2.1 There are two types of particles reacting (left circle), one with two blue atoms and the other with two orange, so the depiction shows a mixture of two elements. In the product (right circle), all the particles have one blue atom and one orange; this is a compound.

2.2 Mass (t) of pitchblende

$$= 2.3 \cancel{\text{ t uranium}} \times \frac{84.2 \text{ t pitchblende}}{71.4 \cancel{\text{ t uranium}}} = 2.7 \text{ t pitchblende}$$

Mass (t) of oxygen

$$= 2.7 \cancel{\text{ t pitchblende}} \times \frac{(84.2 - 71.4 \text{ t oxygen})}{84.2 \cancel{\text{ t pitchblende}}} = 0.41 \text{ t oxygen}$$

2.3 Sample B. Two bromine-fluorine compounds appear. In one, there are three fluorine atoms for each bromine; in the other, there is one fluorine for each bromine. Therefore, in the two compounds, the ratio of fluorines combining with one bromine is 3/1.

2.4 (a) $5p^+$, $6n^0$, $5e^-$; Q = B
(b) $20p^+$, $21n^0$, $20e^-$; R = Ca
(c) $53p^+$, $78n^0$, $53e^-$; X = I

2.5 $10.0129x + [11.0093(1 - x)] = 10.81$; $0.9964x = 0.1993$; $x = 0.2000$ and $1 - x = 0.8000$; % abundance of $^{10}B = 20.00\%$; % abundance of $^{11}B = 80.00\%$

2.6 (a) S^{2-}; (b) Rb^+; (c) Ba^{2+}

2.7 (a) Zinc [Group 2B(12)] and oxygen [Group 6A(16)]
(b) Silver [Group 1B(11)] and bromine [Group 7A(17)]
(c) Lithium [Group 1A(1)] and chlorine [Group 7A(17)]
(d) Aluminum [Group 3A(13)] and sulfur [Group 6A(16)]

2.8 (a) ZnO; (b) AgBr; (c) LiCl; (d) Al_2S_3

2.9 (a) PbO_2; (b) copper(I) sulfide (cuprous sulfide); (c) iron(II) bromide (ferrous bromide); (d) $HgCl_2$

2.10 (a) $Cu(NO_3)_2 \cdot 3H_2O$; (b) $Zn(OH)_2$; (c) lithium cyanide

2.11 (a) $(NH_4)_3PO_4$; ammonium is NH_4^+ and phosphate is PO_4^{3-}.
(b) $Al(OH)_3$; parentheses are needed around the polyatomic ion OH^-.
(c) Magnesium hydrogen carbonate; Mg^{2+} is magnesium and can have only a 2+ charge, so it does not need (II); HCO_3^- is hydrogen carbonate (or bicarbonate).
(d) Chromium(III) nitrate; the *-ic* ending is not used with Roman numerals; NO_3^- is nitr*ate*.
(e) Calcium nitrite; Ca^{2+} is calcium and NO_2^- is nitr*ite*.

2.12 (a) $HClO_3$; (b) hydrofluoric acid; (c) CH_3COOH (or $HC_2H_3O_2$); (d) H_2SO_3; (e) hypobromous acid

2.13 (a) Sulfur trioxide; (b) silicon dioxide; (c) N_2O; (d) SeF_6

2.14 (a) Disulfur dichloride; the *-ous* suffix is not used.
(b) NO; the name indicates one nitrogen.
(c) Bromine trichloride; Br is in a higher period in Group 7A(17), so it is named first.

2.15 (a) H_2O_2, 34.02 amu; (b) CsCl, 168.4 amu; (c) H_2SO_4, 98.09 amu; (d) K_2SO_4, 174.27 amu

2.16 (a) Na_2O. This is an ionic compound, so the name is sodium oxide.

Formula mass

$$= (2 \times \text{atomic mass of Na}) + (1 \times \text{atomic mass of O})$$
$$= (2 \times 22.99 \text{ amu}) + 16.00 \text{ amu} = 61.98 \text{ amu}$$

(b) NO_2. This is a covalent compound, and N has the lower group number, so the name is nitrogen dioxide.

Molecular mass

$$= (1 \times \text{atomic mass of N}) + (2 \times \text{atomic mass of O})$$
$$= 14.01 \text{ amu} + (2 \times 16.00 \text{ amu}) = 46.01 \text{ amu}$$

PROBLEMS

Problems with **colored** numbers are answered in Appendix E and worked in detail in the Student Solutions Manual. Problem sections match those in the text and provide the numbers of relevant sample problems. Most offer Concept Review Questions, Skill-Building Exercises (grouped in pairs covering the same concept), and Problems in Context. The Comprehensive Problems are based on material from any section or previous chapter.

Elements, Compounds, and Mixtures: An Atomic Overview

(Sample Problem 2.1)

Concept Review Questions

• **2.1** What is the key difference between an element and a compound?

• **2.2** List two differences between a compound and a mixture.

• **2.3** Which of the following are pure substances? Explain.
(a) Calcium chloride, used to melt ice on roads, consists of two elements, calcium and chlorine, in a fixed mass ratio.
(b) Sulfur consists of sulfur atoms combined into octatomic molecules.
(c) Baking powder, a leavening agent, contains 26% to 30% sodium hydrogen carbonate and 30% to 35% calcium dihydrogen phosphate by mass.
(d) Cytosine, a component of DNA, consists of H, C, N, and O atoms bonded in a specific arrangement.

• **2.4** Classify each substance in Problem 2.3 as an element, compound, or mixture, and explain your answers.

• **2.5** Explain the following statement: The smallest particles unique to an element may be atoms or molecules.

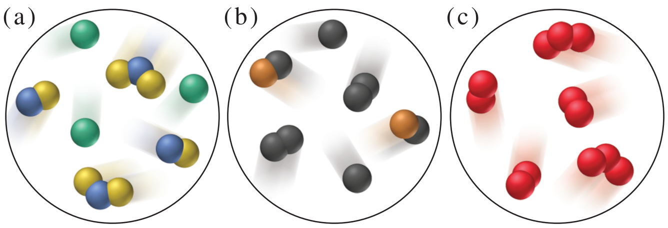
• **2.6** Explain the following statement: The smallest particles unique to a compound cannot be atoms.

• **2.7** Can the relative amounts of the components of a mixture vary? Can the relative amounts of the components of a compound vary? Explain.

Problems in Context

• **2.8** The tap water found in many areas of the United States leaves white deposits when it evaporates. Is this tap water a mixture or a compound? Explain.

• **2.9** Each scene below represents a mixture. Describe each one in terms of the number of elements and/or compounds present.



• **2.10** Samples of illicit "street" drugs often contain an inactive component, such as ascorbic acid (vitamin C). After obtaining a sample of cocaine, government chemists calculate the mass of vitamin C per gram of drug sample, and use it to track the drug's distribution. For example, if different samples of cocaine obtained on the streets of New York, Los Angeles, and Paris all contain 0.6384 g of vitamin C per gram of sample, they very likely come from a common source. Do these street samples consist of a compound, element, or mixture? Explain.

The Observations That Led to an Atomic View of Matter

(Sample Problem 2.2)

Concept Review Questions

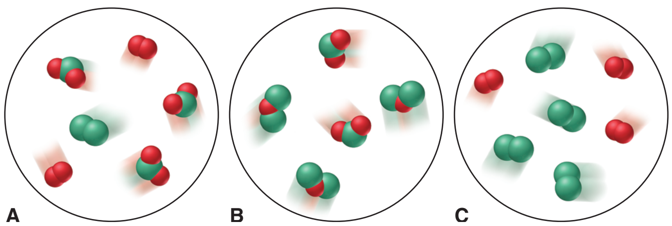
• **2.11** Why was it necessary for separation techniques and methods of chemical analysis to be developed before the laws of definite composition and multiple proportions could be formulated?

• **2.12** To which classes of matter—element, compound, and/or mixture—do the following apply: (a) law of mass conservation; (b) law of definite composition; (c) law of multiple proportions?

• **2.13** In our modern view of matter and energy, is the law of mass conservation still relevant to chemical reactions? Explain.

• **2.14** Identify the mass law that each of the following observations demonstrates, and explain your reasoning:
(a) A sample of potassium chloride from Chile contains the same percent by mass of potassium as one from Poland.
(b) A flashbulb contains magnesium and oxygen before use and magnesium oxide afterward, but its mass does not change.
(c) Arsenic and oxygen form one compound that is 65.2 mass % arsenic and another that is 75.8 mass % arsenic.

• **2.15** Which of the following scenes illustrate(s) the fact that compounds of chlorine (green) and oxygen (red) exhibit the law of multiple proportions? Name the compounds.



•• **2.16** (a) Does the percent by mass of each element in a compound depend on the amount of compound? Explain. (b) Does the mass of each element in a compound depend on the amount of compound? Explain.

• **2.17** Does the percent by mass of each element in a compound depend on the amount of that element used to make the compound? Explain.

Skill-Building Exercises *(grouped in similar pairs)*

• **2.18** State the mass law(s) demonstrated by the following experimental results, and explain your reasoning:
Experiment 1: A student heats 1.00 g of a blue compound and obtains 0.64 g of a white compound and 0.36 g of a colorless gas.
Experiment 2: A second student heats 3.25 g of the same blue compound and obtains 2.08 g of a white compound and 1.17 g of a colorless gas.

• **2.19** State the mass law(s) demonstrated by the following experimental results, and explain your reasoning:
Experiment 1: A student heats 1.27 g of copper and 3.50 g of iodine to produce 3.81 g of a white compound; 0.96 g of iodine remains.
Experiment 2: A second student heats 2.55 g of copper and 3.50 g of iodine to form 5.25 g of a white compound, and 0.80 g of copper remains.

• **2.20** Fluorite, a mineral of calcium, is a compound of the metal with fluorine. Analysis shows that a 2.76-g sample of fluorite contains 1.42 g of calcium. Calculate the (a) mass of fluorine in the sample; (b) mass fractions of calcium and fluorine in fluorite; (c) mass percents of calcium and fluorine in fluorite.

• **2.21** Galena, a mineral of lead, is a compound of the metal with sulfur. Analysis shows that a 2.34-g sample of galena contains 2.03 g of lead. Calculate the (a) mass of sulfur in the sample; (b) mass fractions of lead and sulfur in galena; (c) mass percents of lead and sulfur in galena.

• **2.22** Magnesium oxide (MgO) forms when the metal burns in air. (a) If 1.25 g of MgO contains 0.754 g of Mg, what is the mass ratio of magnesium to oxide? (b) How many grams of Mg are in 534 g of MgO?

• **2.23** Zinc sulfide (ZnS) occurs in the zinc blende crystal structure. (a) If 2.54 g of ZnS contains 1.70 g of Zn, what is the mass ratio of zinc to sulfide? (b) How many kilograms of Zn are in 3.82 kg of ZnS?

• **2.24** A compound of copper and sulfur contains 88.39 g of metal and 44.61 g of nonmetal. How many grams of copper are in 5264 kg of compound? How many grams of sulfur?

• **2.25** A compound of iodine and cesium contains 63.94 g of metal and 61.06 g of nonmetal. How many grams of cesium are in 38.77 g of compound? How many grams of iodine?

• **2.26** Show, with calculations, how the following data illustrate the law of multiple proportions:
Compound 1: 47.5 mass % sulfur and 52.5 mass % chlorine
Compound 2: 31.1 mass % sulfur and 68.9 mass % chlorine

• **2.27** Show, with calculations, how the following data illustrate the law of multiple proportions:
Compound 1: 77.6 mass % xenon and 22.4 mass % fluorine
Compound 2: 63.3 mass % xenon and 36.7 mass % fluorine

Problems in Context

• **2.28** Dolomite is a carbonate of magnesium and calcium. Analysis shows that 7.81 g of dolomite contains 1.70 g of Ca. Calculate the mass percent of Ca in dolomite. On the basis of the mass percent of Ca, and neglecting all other factors, which is the richer source of Ca, dolomite or fluorite (see Problem 2.20)?

• **2.29** The mass percent of sulfur in a sample of coal is a key factor in the environmental impact of the coal because the sulfur combines with oxygen when the coal is burned and the oxide can then be incorporated into acid rain. Which of the following coals would have the smallest environmental impact?

	Mass (g) of Sample	Mass (g) of Sulfur in Sample
Coal A	378	11.3
Coal B	495	19.0
Coal C	675	20.6

Dalton's Atomic Theory

(Sample Problem 2.3)

Concept Review Questions

• **2.30** Which of Dalton's postulates about atoms are inconsistent with later observations? Do these inconsistencies mean that Dalton was wrong? Is Dalton's model still useful? Explain.

• **2.31** Use Dalton's theory to explain why potassium nitrate from India or Italy has the same mass percents of K, N, and O.

The Observations That Led to the Nuclear Atom Model

Concept Review Questions

• **2.32** Thomson was able to determine the mass/charge ratio of the electron but not its mass. How did Millikan's experiment allow determination of the electron's mass?

•• **2.33** The following charges on individual oil droplets were obtained during an experiment similar to Millikan's. Determine a charge for the electron (in C, coulombs), and explain your answer: -3.204×10^{-19} C; -4.806×10^{-19} C; -8.010×10^{-19} C; -1.442×10^{-18} C.

•• **2.34** Describe Thomson's model of the atom. How might it account for the production of cathode rays?

• **2.35** When Rutherford's coworkers bombarded gold foil with α particles, they obtained results that overturned the existing (Thomson) model of the atom. Explain.

The Atomic Theory Today

(Sample Problems 2.4 and 2.5)

Concept Review Questions

• **2.36** Define *atomic number* and *mass number.* Which can vary without changing the identity of the element?

• **2.37** Choose the correct answer. The difference between the mass number of an isotope and its atomic number is (a) directly related to the identity of the element; (b) the number of electrons; (c) the number of neutrons; (d) the number of isotopes.

• **2.38** Even though several elements have only one naturally occurring isotope and all atomic nuclei have whole numbers of protons and neutrons, no atomic mass is a whole number. Use the data from Table 2.2 to explain this fact.

Skill-Building Exercises *(grouped in similar pairs)*

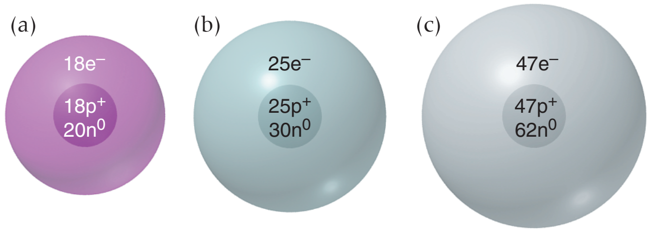
• **2.39** Argon has three naturally occurring isotopes, ${}^{36}Ar$, ${}^{38}Ar$, and ${}^{40}Ar$. What is the mass number of each? How many protons, neutrons, and electrons are present in each?

• **2.40** Chlorine has two naturally occurring isotopes, ${}^{35}Cl$ and ${}^{37}Cl$. What is the mass number of each isotope? How many protons, neutrons, and electrons are present in each?

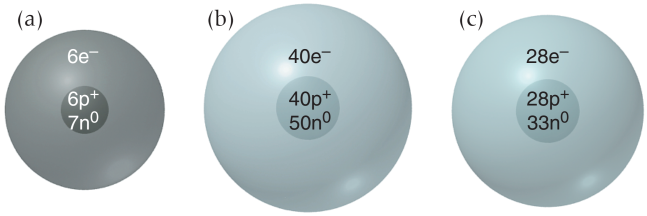
•• **2.41** Do both members of the following pairs have the same number of protons? Neutrons? Electrons?
(a) ${}^{16}_{8}O$ and ${}^{17}_{8}O$ (b) ${}^{40}_{18}Ar$ and ${}^{41}_{19}K$ (c) ${}^{60}_{27}Co$ and ${}^{60}_{28}Ni$
Which pair(s) consist(s) of atoms with the same *Z* value? *N* value? *A* value?

•• **2.42** Do both members of the following pairs have the same number of protons? Neutrons? Electrons?
(a) ${}^{3}_{1}H$ and ${}^{3}_{2}He$ (b) ${}^{14}_{6}C$ and ${}^{15}_{7}N$ (c) ${}^{19}_{9}F$ and ${}^{18}_{9}F$
Which pair(s) consist(s) of atoms with the same *Z* value? *N* value? *A* value?

• **2.43** Write the ${}^{A}_{Z}X$ notation for each atomic depiction:



• **2.44** Write the $^A_Z X$ notation for each atomic depiction:



• **2.45** Draw atomic depictions similar to those in Problem 2.43 for (a) $^{48}_{22}$Ti; (b) $^{79}_{34}$Se; (c) $^{11}_{5}$B.

• **2.46** Draw atomic depictions similar to those in Problem 2.43 for (a) $^{207}_{82}$Pb; (b) $^{9}_{4}$Be; (c) $^{75}_{33}$As.

• **2.47** Gallium has two naturally occurring isotopes, ^{69}Ga (isotopic mass 68.9256 amu, abundance 60.11%) and ^{71}Ga (isotopic mass 70.9247 amu, abundance 39.89%). Calculate the atomic mass of gallium.

• **2.48** Magnesium has three naturally occurring isotopes, ^{24}Mg (isotopic mass 23.9850 amu, abundance 78.99%), ^{25}Mg (isotopic mass 24.9858 amu, abundance 10.00%), and ^{26}Mg (isotopic mass 25.9826 amu, abundance 11.01%). Calculate the atomic mass of magnesium.

•• **2.49** Chlorine has two naturally occurring isotopes, ^{35}Cl (isotopic mass 34.9689 amu) and ^{37}Cl (isotopic mass 36.9659 amu). If chlorine has an atomic mass of 35.4527 amu, what is the percent abundance of each isotope?

•• **2.50** Copper has two naturally occurring isotopes, ^{63}Cu (isotopic mass 62.9396 amu) and ^{65}Cu (isotopic mass 64.9278 amu). If copper has an atomic mass of 63.546 amu, what is the percent abundance of each isotope?

Elements: A First Look at the Periodic Table

● Concept Review Questions

•• **2.51** How can iodine ($Z = 53$) have a higher atomic number yet a lower atomic mass than tellurium ($Z = 52$)?

• **2.52** Correct each of the following statements:
(a) In the modern periodic table, the elements are arranged in order of increasing atomic mass.
(b) Elements in a period have similar chemical properties.
(c) Elements can be classified as either metalloids or nonmetals.

• **2.53** What class of elements lies along the "staircase" line in the periodic table? How do their properties compare with those of metals and nonmetals?

• **2.54** What are some characteristic properties of elements to the left of the elements along the "staircase"? To the right?

• **2.55** The elements in Groups 1A(1) and 7A(17) are all quite reactive. What is a major difference between them?

● Skill-Building Exercises *(grouped in similar pairs)*

• **2.56** Give the name, atomic symbol, and group number of the element with the following Z value, and classify it as a metal, metalloid, or nonmetal:
(a) $Z = 32$ (b) $Z = 15$ (c) $Z = 2$ (d) $Z = 3$ (e) $Z = 42$

• **2.57** Give the name, atomic symbol, and group number of the element with the following Z value, and classify it as a metal, metalloid, or nonmetal:
(a) $Z = 33$ (b) $Z = 20$ (c) $Z = 35$ (d) $Z = 19$ (e) $Z = 13$

• **2.58** Fill in the blanks:
(a) The symbol and atomic number of the heaviest alkaline earth metal are ____________ and ____________.
(b) The symbol and atomic number of the lightest metalloid in Group 4A(14) are ____________ and ____________.
(c) Group 1B(11) consists of the *coinage metals*. The symbol and atomic mass of the coinage metal whose atoms have the fewest electrons are ____________ and ____________.
(d) The symbol and atomic mass of the halogen in Period 4 are ____________ and ____________.

• **2.59** Fill in the blanks:
(a) The symbol and atomic number of the heaviest nonradioactive noble gas are ____________ and ____________.
(b) The symbol and group number of the Period 5 transition element whose atoms have the fewest protons are ____________ and ____________.
(c) The elements in Group 6A(16) are sometimes called the *chalcogens*. The symbol and atomic number of the first metallic chalcogen are ____________ and ____________.
(d) The symbol and number of protons of the Period 4 alkali metal atom are ____________ and ____________.

Compounds: Introduction to Bonding

(Sample Problem 2.6)

● Concept Review Questions

• **2.60** Describe the type and nature of the bonding that occurs between reactive metals and nonmetals.

• **2.61** Describe the type and nature of the bonding that often occurs between two nonmetals.

• **2.62** How can ionic compounds be neutral if they consist of positive and negative ions?

• **2.63** Given that the ions in LiF and in MgO are of similar size, which compound has stronger ionic bonding? Use Coulomb's law in your explanation.

• **2.64** Are molecules present in a sample of BaF_2? Explain.

• **2.65** Are ions present in a sample of P_4O_6? Explain.

• **2.66** The monatomic ions of Groups 1A(1) and 7A(17) are all singly charged. In what major way do they differ? Why?

• **2.67** Describe the formation of solid magnesium chloride ($MgCl_2$) from large numbers of magnesium and chlorine atoms.

• **2.68** Describe the formation of solid potassium sulfide (K_2S) from large numbers of potassium and sulfur atoms.

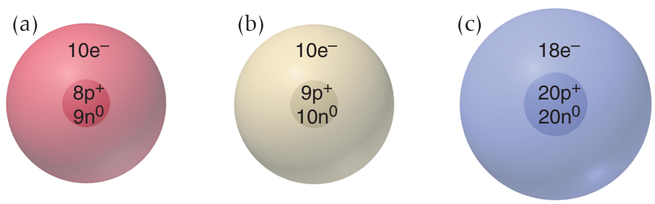
• **2.69** Does potassium nitrate (KNO_3) incorporate ionic bonding, covalent bonding, or both? Explain.

● Skill-Building Exercises *(grouped in similar pairs)*

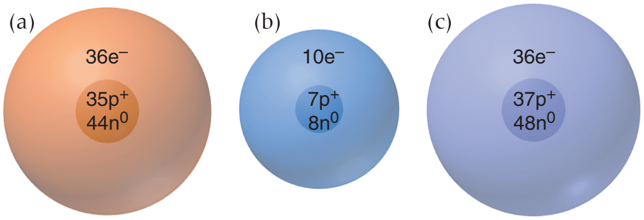
• **2.70** What monatomic ions do potassium ($Z = 19$) and iodine ($Z = 53$) form?

• **2.71** What monatomic ions do barium ($Z = 56$) and selenium ($Z = 34$) form?

• **2.72** For each ionic depiction, give the name of the parent atom, its mass number, and its group and period numbers:



• **2.73** For each ionic depiction, give the name of the parent atom, its mass number, and its group and period numbers:



•• **2.74** An ionic compound forms when lithium ($Z = 3$) reacts with oxygen ($Z = 8$). If a sample of the compound contains 8.4×10^{21} lithium ions, how many oxide ions does it contain?

•• **2.75** An ionic compound forms when calcium ($Z = 20$) reacts with iodine ($Z = 53$). If a sample of the compound contains 7.4×10^{21} calcium ions, how many iodide ions does it contain?

• **2.76** The radii of the sodium and potassium ions are 102 pm and 138 pm, respectively. Which compound has stronger ionic attractions, sodium chloride or potassium chloride?

• **2.77** The radii of the lithium and magnesium ions are 76 pm and 72 pm, respectively. Which compound has stronger ionic attractions, lithium oxide or magnesium oxide?

Compounds: Formulas, Names, and Masses

(Sample Problems 2.7 to 2.16)

Concept Review Questions

• **2.78** What is the difference between an empirical formula and a molecular formula? Can they ever be the same?

• **2.79** How is a structural formula similar to a molecular formula? How is it different?

•• **2.80** Consider a mixture of 10 billion O_2 molecules and 10 billion H_2 molecules. In what way is this mixture similar to a sample containing 10 billion hydrogen peroxide (H_2O_2) molecules? In what way is it different?

• **2.81** For what type(s) of compound do we use Roman numerals in the name?

• **2.82** For what type(s) of compound do we use Greek numerical prefixes in the name?

• **2.83** For what type of compound are we unable to write a molecular formula?

Skill-Building Exercises *(grouped in similar pairs)*

• **2.84** Write an empirical formula for each of the following:
(a) Hydrazine, a rocket fuel, molecular formula N_2H_4
(b) Glucose, a sugar, molecular formula $C_6H_{12}O_6$

• **2.85** Write an empirical formula for each of the following:
(a) Ethylene glycol, car antifreeze, molecular formula $C_2H_6O_2$
(b) Peroxodisulfuric acid, a compound used to make bleaching agents, molecular formula $H_2S_2O_8$

•• **2.86** Give the name and formula of the compound formed from the following elements: (a) sodium and nitrogen; (b) oxygen and strontium; (c) aluminum and chlorine.

•• **2.87** Give the name and formula of the compound formed from the following elements: (a) cesium and bromine; (b) sulfur and barium; (c) calcium and fluorine.

•• **2.88** Give the name and formula of the compound formed from the following elements:
(a) ${}_{12}L$ and ${}_{9}M$ (b) ${}_{30}L$ and ${}_{16}M$ (c) ${}_{17}L$ and ${}_{38}M$

•• **2.89** Give the name and formula of the compound formed from the following elements:
(a) ${}_{37}Q$ and ${}_{35}R$ (b) ${}_{8}Q$ and ${}_{13}R$ (c) ${}_{20}Q$ and ${}_{53}R$

• **2.90** Give the systematic names for the formulas or the formulas for the names: (a) tin(IV) chloride; (b) $FeBr_3$; (c) cuprous bromide; (d) Mn_2O_3.

• **2.91** Give the systematic names for the formulas or the formulas for the names: (a) Na_2HPO_4; (b) potassium carbonate dihydrate; (c) $NaNO_2$; (d) ammonium perchlorate.

• **2.92** Give the systematic names for the formulas or the formulas for the names: (a) CoO; (b) mercury(I) chloride; (c) $Pb(C_2H_3O_2)_2 \cdot 3H_2O$; (d) chromic oxide.

• **2.93** Give the systematic names for the formulas or the formulas for the names: (a) $Sn(SO_3)_2$; (b) potassium dichromate; (c) $FeCO_3$; (d) copper(II) nitrate.

• **2.94** Correct each of the following formulas:
(a) Barium oxide is BaO_2.
(b) Iron(II) nitrate is $Fe(NO_3)_3$.
(c) Magnesium sulfide is $MnSO_3$.

• **2.95** Correct each of the following names:
(a) CuI is cobalt(II) iodide.
(b) $Fe(HSO_4)_3$ is iron(II) sulfate.
(c) $MgCr_2O_7$ is magnesium dichromium heptaoxide.

• **2.96** Give the name and formula for the acid derived from each of the following anions:
(a) hydrogen sulfate (b) IO_3^- (c) cyanide (d) HS^-

• **2.97** Give the name and formula for the acid derived from each of the following anions:
(a) perchlorate (b) NO_3^- (c) bromite (d) F^-

• **2.98** Many chemical names are similar at first glance. Give the formulas of the species in each set: (a) ammonium ion and ammonia; (b) magnesium sulfide, magnesium sulfite, and magnesium sulfate; (c) hydrochloric acid, chloric acid, and chlorous acid; (d) cuprous bromide and cupric bromide.

• **2.99** Give the formulas of the compounds in each set: (a) lead(II) oxide and lead(IV) oxide; (b) lithium nitride, lithium nitrite, and lithium nitrate; (c) strontium hydride and strontium hydroxide; (d) magnesium oxide and manganese(II) oxide.

• **2.100** Give the name and formula of the compound whose molecules consist of two sulfur atoms and four fluorine atoms.

• **2.101** Give the name and formula of the compound whose molecules consist of two chlorine atoms and one oxygen atom.

• **2.102** Correct the name to match the formula of the following compounds: (a) calcium(II) dichloride, $CaCl_2$; (b) copper(II) oxide, Cu_2O; (c) stannous tetrafluoride, SnF_4; (d) hydrogen chloride acid, HCl.

• **2.103** Correct the formula to match the name of the following compounds: (a) iron(III) oxide, Fe_3O_4; (b) chloric acid, HCl; (c) mercuric oxide, Hg_2O; (d) dichlorine heptaoxide, Cl_2O_6.

•• **2.104** Give the number of atoms of the specified element in a formula unit of each of the following compounds, and calculate the molecular (formula) mass:
(a) Oxygen in aluminum sulfate, $Al_2(SO_4)_3$
(b) Hydrogen in ammonium hydrogen phosphate, $(NH_4)_2HPO_4$
(c) Oxygen in the mineral azurite, $Cu_3(OH)_2(CO_3)_2$

•• **2.105** Give the number of atoms of the specified element in a formula unit of each of the following compounds, and calculate the molecular (formula) mass:
(a) Hydrogen in ammonium benzoate, $C_6H_5COONH_4$
(b) Nitrogen in hydrazinium sulfate, $N_2H_6SO_4$
(c) Oxygen in the mineral leadhillite, $Pb_4SO_4(CO_3)_2(OH)_2$

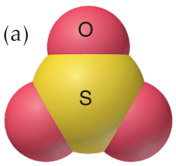
•• **2.106** Write the formula of each compound, and determine its molecular (formula) mass: (a) ammonium sulfate; (b) sodium dihydrogen phosphate; (c) potassium bicarbonate.

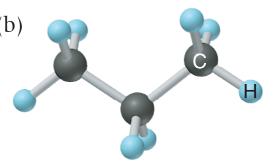
•• **2.107** Write the formula of each compound, and determine its molecular (formula) mass: (a) sodium dichromate; (b) ammonium perchlorate; (c) magnesium nitrite trihydrate.

•• **2.108** Calculate the molecular (formula) mass of each compound: (a) dinitrogen pentaoxide; (b) lead(II) nitrate; (c) calcium peroxide.

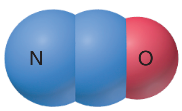
•• **2.109** Calculate the molecular (formula) mass of each compound: (a) iron(II) acetate tetrahydrate; (b) sulfur tetrachloride; (c) potassium permanganate.

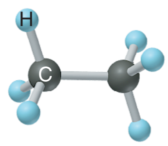
•• **2.110** Give the formula, name, and molecular mass of the following molecules:

(a)


(b)


•• **2.111** Give the formula, name, and molecular mass of the following molecules:

(a)


(b)


•• **2.112** Give the name, empirical formula, and molecular mass of the molecule depicted in Figure P2.112.

•• **2.113** Give the name, empirical formula, and molecular mass of the molecule depicted in Figure P2.113.

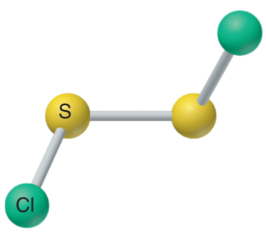


Figure P2.112

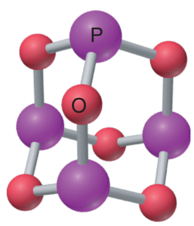
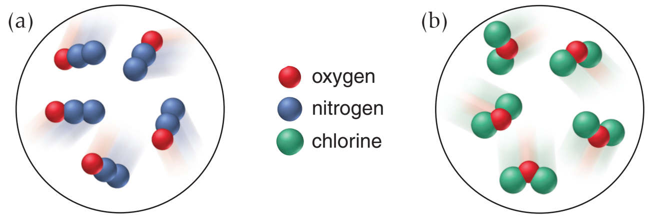


Figure P2.113

Problems in Context

• **2.114** Before the use of systematic names, many compounds had common names. Give the systematic name for each of the following: (a) blue vitriol, $CuSO_4{\cdot}5H_2O$; (b) slaked lime, $Ca(OH)_2$; (c) oil of vitriol, H_2SO_4; (d) washing soda, Na_2CO_3; (e) muriatic acid, HCl; (f) Epsom salt, $MgSO_4{\cdot}7H_2O$; (g) chalk, $CaCO_3$; (h) dry ice, CO_2; (i) baking soda, $NaHCO_3$; (j) lye, NaOH.

•• **2.115** Each circle contains a representation of a binary compound. Determine its name, formula, and molecular (formula) mass.



Mixtures: Classification and Separation

Concept Review Questions

• **2.116** In what main way is separating the components of a mixture different from separating the components of a compound?

• **2.117** What is the difference between a homogeneous and a heterogeneous mixture?

• **2.118** Is a solution a homogeneous or a heterogeneous mixture? Give an example of an aqueous solution.

Skill-Building Exercises *(grouped in similar pairs)*

• **2.119** Classify each of the following as a compound, a homogeneous mixture, or a heterogeneous mixture: (a) distilled water; (b) gasoline; (c) beach sand; (d) wine; (e) air.

• **2.120** Classify each of the following as a compound, a homogeneous mixture, or a heterogeneous mixture: (a) orange juice; (b) vegetable soup; (c) cement; (d) calcium sulfate; (e) tea.

•• **2.121** Name the technique(s) and briefly describe the procedure you would use to separate each of the following mixtures into two components: (a) table salt and pepper; (b) table sugar and sand; (c) drinking water contaminated with fuel oil; (d) vegetable oil and vinegar.

•• **2.122** Name the technique(s) and briefly describe the procedure you would use to separate each of the following mixtures into two components: (a) crushed ice and crushed glass; (b) table sugar dissolved in ethanol; (c) iron and sulfur; (d) two pigments (chlorophyll *a* and chlorophyll *b*) from spinach leaves.

Problems in Context

•• **2.123** Which separation method is operating in each of the following procedures: (a) pouring a mixture of cooked pasta and boiling water into a colander; (b) removing colored impurities from raw sugar to make refined sugar; (c) preparing coffee by pouring hot water through ground coffee beans?

•• **2.124** A quality-control laboratory analyzes a product mixture using gas-liquid chromatography. The separation of components is more than adequate, but the process takes too long. Suggest two ways, other than changing the stationary phase, to shorten the analysis time.

Comprehensive Problems

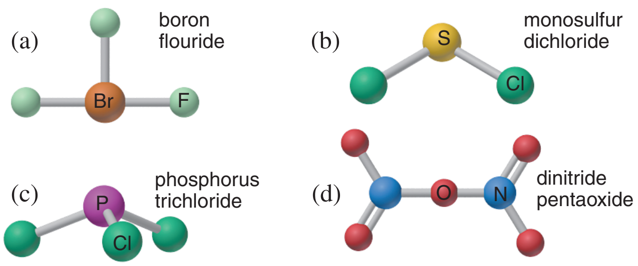
•• **2.125** Helium is the lightest noble gas and the second most abundant element (after hydrogen) in the universe.
(a) The radius of a helium atom is 3.1×10^{-11} m; the radius of its nucleus is 2.5×10^{-15} m. What fraction of the spherical atomic volume is occupied by the nucleus (V of a sphere $= \frac{4}{3}\pi r^3$)?
(b) The mass of a helium-4 atom is 6.64648×10^{-24} g, and each of its two electrons has a mass of 9.10939×10^{-28} g. What fraction of this atom's mass is contributed by its nucleus?

This figure is available as a transparency and is also on the Presentation Center, within the art library.

•• **2.126** From the following ions and their radii (in pm), choose a pair that gives the strongest ionic bonding and a pair that gives the weakest: Mg^{2+} 72; K^+ 138; Rb^+ 152; Ba^{2+} 135; Cl^- 181; O^{2-} 140; I^- 220.

•• **2.127** Prior to 1961, the atomic mass standard was defined as $\frac{1}{16}$ of the mass of ^{16}O. Based on that standard: (a) What was the mass of carbon-12, given the modern atomic mass of oxygen is 15.9994 amu? (b) What was the mass of potassium-39, given its modern isotopic mass is 38.9637 amu?

• **2.128** Give the molecular mass of each compound depicted below, and provide a correct name for any that are named incorrectly.



•• **2.129** Transition metals, located in the center of the periodic table, have many essential uses as elements and form many important compounds as well. Calculate the molecular mass of the following transition metal compounds:
(a) $[Co(NH_3)_6]Cl_3$ (b) $[Pt(NH_3)_4BrCl]Cl_2$
(c) $K_4[V(CN)_6]$ (d) $[Ce(NH_3)_6][FeCl_4]_3$

••• **2.130** A rock is 5.0% by mass fayalite (Fe_2SiO_4), 7.0% by mass forsterite (Mg_2SiO_4), and the remainder silicon dioxide. What is the mass percent of each element in the rock?

•• **2.131** Polyatomic ions are named by patterns that apply to elements in a given group. Using the periodic table and Table 2.5, give the name of each of the following: (a) SeO_4^{2-}; (b) AsO_4^{3-}; (c) BrO_2^-; (d) $HSeO_4^-$; (e) TeO_3^{2-}.

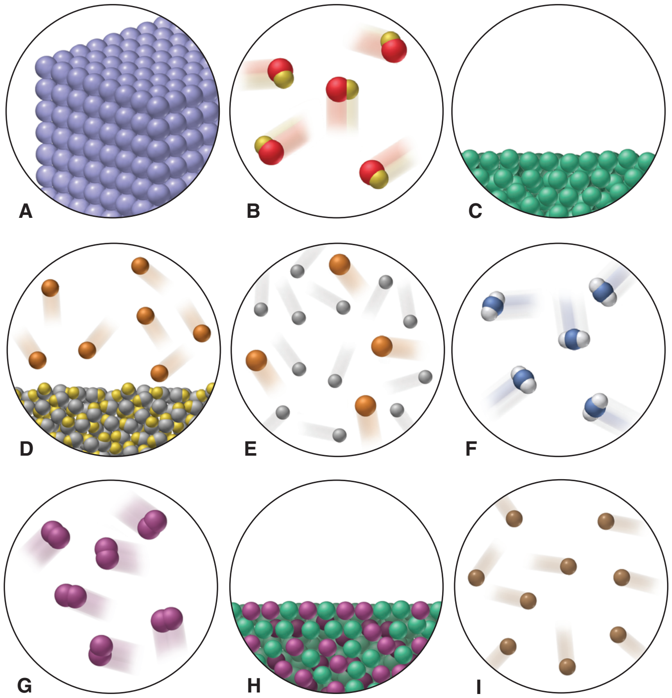
•• **2.132** Ammonium dihydrogen phosphate, formed from the reaction of phosphoric acid with ammonia, is used as a crop fertilizer as well as a component of some fire extinguishers. (a) What are the mass percentages of N and P in the compound? (b) How much ammonia is incorporated into 100. g of compound?

•• **2.133** Nitrogen forms more oxides than any other element. The percents by mass of N in three different nitrogen oxides are (I) 46.69%; (II) 36.85%; (III) 25.94%. (a) Determine the empirical formula of each compound. (b) How many grams of oxygen per 1.00 g of nitrogen are in each compound?

••• **2.134** Boron has two naturally occurring isotopes, ^{10}B (19.9%) and ^{11}B (80.1%). Although the B_2 molecule does not exist naturally on Earth, it has been produced in the laboratory and been observed in stars. (a) How many different B_2 molecules are possible? (b) What are the masses and percent abundances of each?

•• **2.135** Dimercaprol ($HSCH_2CHSHCH_2OH$) was developed during World War I as an antidote to arsenic-based poison gas and is used today to treat heavy-metal poisoning. It binds the toxic element and carries it out of the body.
(a) If each molecule binds one arsenic (As) atom, how many atoms of As could be removed by 175 mg of dimercaprol?
(b) If one molecule binds one metal atom, calculate the mass % of each of the following metals in a metal-dimercaprol combination: mercury, thallium, chromium.

•• **2.136** Scenes A–I depict various types of matter on the atomic scale. Choose the correct scene(s) for each of the following:
(a) A mixture that fills its container
(b) A substance that cannot be broken down into simpler ones
(c) An element with a very high resistance to flow
(d) A homogeneous mixture
(e) An element that conforms to the walls of its container and displays a surface
(f) A gas consisting of diatomic particles
(g) A gas that can be broken down into simpler substances
(h) A substance with a 2/1 ratio of its component atoms
(i) Matter that can be separated into its component substances by physical means
(j) A heterogeneous mixture
(k) Matter that obeys the law of definite composition



•• **2.137** The number of atoms in 1 dm^3 of aluminum is nearly the same as the number of atoms in 1 dm^3 of lead, but the densities of these metals are very different (see Table 1.5). Explain.

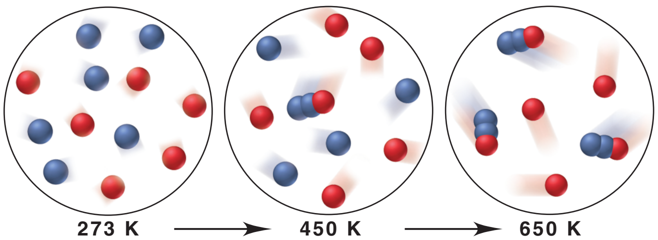
•• **2.138** You are working in the laboratory preparing sodium chloride. Consider the following results for three preparations of the compound:
Case 1: 39.34 g Na + 60.66 g $Cl_2 \longrightarrow$ 100.00 g NaCl
Case 2: 39.34 g Na + 70.00 g $Cl_2 \longrightarrow$
100.00 g NaCl + 9.34 g Cl_2
Case 3: 50.00 g Na + 50.00 g $Cl_2 \longrightarrow$
82.43 g NaCl + 17.57 g Na
Explain these results in terms of the laws of conservation of mass and definite composition.

•• **2.139** The seven most abundant ions in seawater make up more than 99% by mass of the dissolved compounds. They are listed in units of mg ion/kg seawater: chloride 18,980; sodium 10,560; sulfate 2650; magnesium 1270; calcium 400; potassium 380; hydrogen carbonate 140.
(a) What is the mass % of each ion in seawater?
(b) What percent of the total mass of ions is sodium ion?

Environmental: 2.139

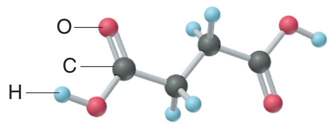
(c) How does the total mass % of alkaline earth metal ions compare with the total mass % of alkali metal ions?
(d) Which makes up the larger mass fraction of dissolved components, anions or cations?

•• **2.140** The scenes below represent a mixture of two monatomic gases undergoing a reaction when heated. Which mass law(s) is (are) illustrated by this change?



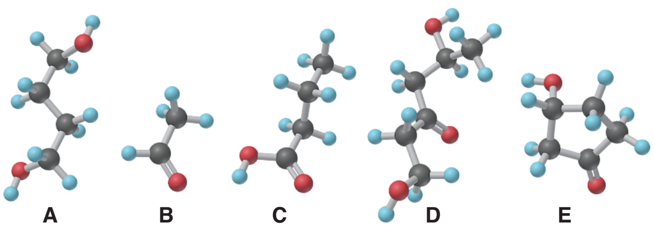
••• **2.141** When barium (Ba) reacts with sulfur (S) to form barium sulfide (BaS), each Ba atom reacts with an S atom. If 2.50 cm^3 of Ba reacts with 1.75 cm^3 of S, are there enough Ba atoms to react with the S atoms (d of Ba = 3.51 g/cm^3; d of S = 2.07 g/cm^3)?

•• **2.142** Succinic acid *(below)* is an important metabolite in biological energy production. Give the molecular formula, empirical formula, and molecular mass of succinic acid, and calculate the mass percent of each element.



•• **2.143** Fluoride ion is poisonous in relatively low amounts: 0.2 g of F^- per 70 kg of body weight can cause death. Nevertheless, in order to prevent tooth decay, F^- ions are added to drinking water at a concentration of 1 mg of F^- ion per L of water. How many liters of fluoridated drinking water would a 70-kg person have to consume in one day to reach this toxic level? How many kilograms of sodium fluoride would be needed to treat a 8.50×10^7-gal reservoir?

•• **2.144** Which of the following models represent compounds having the same empirical formula? What is the molecular mass of this common empirical formula?

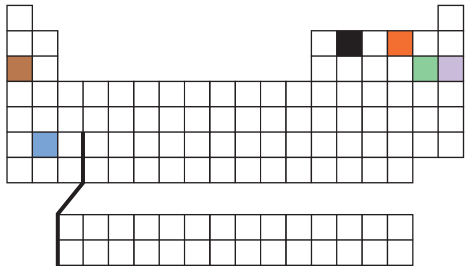


•• **2.145** Antimony has many uses, for example, in infrared devices and as part of an alloy in lead storage batteries. The element has two naturally occurring isotopes, one with mass 120.904 amu, the other with mass 122.904 amu. (a) Write the ^A_ZX notation for each isotope. (b) Use the atomic mass of antimony from the periodic table to calculate the natural abundance of each isotope.

•• **2.146** Dinitrogen monoxide (N_2O; nitrous oxide) is a greenhouse gas that enters the atmosphere principally from natural fertilizer breakdown. Some studies have shown that the isotope ratios of ^{15}N to ^{14}N and of ^{18}O to ^{16}O in N_2O depend on the source, which can thus be determined by measuring the relative abundance of molecular masses in a sample of N_2O.
(a) What different molecular masses are possible for N_2O?
(b) The percent abundance of ^{14}N is 99.6%, and that of ^{16}O is 99.8%. Which molecular mass of N_2O is least common, and which is most common?

•• **2.147** Silver acetylide (AgC_2H) is a shock-sensitive explosive. The synthesis of an organic compound in the presence of silver salts leaves a residue whose mass spectrum shows an ion with m/e 132.9 but no other ions between m/e 130 and 135. Should the chemist be concerned that the residue may be explosive?

•• **2.148** Choose the box color(s) in the periodic table below that match(es) each of the following:



(a) Four elements that are nonmetals
(b) Two elements that are metals
(c) Three elements that are gases at room temperature
(d) Three elements that are solid at room temperature
(e) One pair of elements likely to form a covalent compound
(f) Another pair of elements likely to form a covalent compound
(g) One pair of elements likely to form an ionic compound with formula MX
(h) Another pair of elements likely to form an ionic compound with formula MX
(i) Two elements likely to form an ionic compound with formula M_2X
(j) Two elements likely to form an ionic compound with formula MX_2
(k) An element that forms no compounds
(l) A pair of elements whose compounds exhibit the law of multiple proportions
(m) Two elements that are building blocks in biomolecules
(n) Two elements that are macronutrients in organisms

•• **2.149** The two isotopes of potassium with significant abundance in nature are ^{39}K (isotopic mass 38.9637 amu, 93.258%) and ^{41}K (isotopic mass 40.9618 amu, 6.730%). Fluorine has only one naturally occurring isotope, ^{19}F (isotopic mass 18.9984 amu). Calculate the formula mass of potassium fluoride.

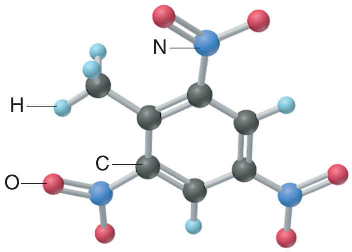
•• **2.150** Boron trifluoride is used as a catalyst in the synthesis of organic compounds. When this compound is analyzed by mass spectrometry (see Tools of the Laboratory, p. 55), several different 1+ ions form, including ions representing the whole molecule as well as molecular fragments formed by the loss of one, two, and three F atoms. Given that boron has two naturally occurring isotopes, ^{10}B and ^{11}B, and fluorine has one, ^{19}F, calculate the masses of all possible 1+ ions.

•• **2.151** Nitrogen monoxide (NO) is a bioactive molecule in blood. Low NO concentrations cause respiratory distress and the formation of blood clots. Doctors prescribe nitroglycerin,

Biological: 2.142, 2.151; Engineering: 2.143, 2.145; Environmental: 2.146; Organic: 2.144, 2.150

$C_3H_5N_3O_9$, and isoamyl nitrate, $(CH_3)_2CHCH_2CH_2ONO_2$, to increase NO. If each compound releases one molecule of NO per atom of N, calculate the mass percent of NO in each medicine.

••• **2.152** TNT (trinitrotoluene; *below*) is used as an explosive in construction. Calculate the mass of each element in 1.00 lb of TNT.



••• **2.153** Carboxylic acids react with alcohols to form esters, which are found in all plants and animals. Some are responsible for the flavors and odors of fruits and flowers. What is the percent by mass of carbon in each of the following esters?

Name	Formula	Odor
Isoamyl isovalerate	$C_4H_9COOC_5H_{11}$	apple
Amyl butyrate	$C_3H_7COOC_5H_{11}$	apricot
Isoamyl acetate	$CH_3COOC_5H_{11}$	banana
Ethyl butyrate	$C_3H_7COOC_2H_5$	pineapple

•• **2.154** Nuclei differ in their stability, and some are so unstable that they undergo radioactive decay. The ratio of the number of neutrons to number of protons (N/Z) in a nucleus correlates with its stability. Calculate the N/Z ratio for (a) ^{144}Sm; (b) ^{56}Fe; (c) ^{20}Ne; (d) ^{107}Ag. (e) The radioactive isotope ^{238}U decays in a series of nuclear reactions that includes another uranium isotope, ^{234}U, and three lead isotopes, ^{214}Pb, ^{210}Pb, and ^{206}Pb. How many neutrons, protons, and electrons are in each of these five isotopes?

•• **2.155** The anticancer drug Platinol (Cisplatin), $Pt(NH_3)_2Cl_2$, reacts with the cancer cell's DNA and interferes with its growth. (a) What is the mass % of platinum (Pt) in Platinol? (b) If Pt costs \$19/g, how many grams of Platinol can be made for \$1.00 million (assume that the cost of Pt determines the cost of the drug)?

•• **2.156** Grignard reagents, which contain a C—Mg bond, have the general formula $CH_3—(CH_2)_x—MgBr$ and are essential in the synthesis of organic compounds. (a) Calculate the mass percent of Mg if $x = 0$. (b) Calculate the mass percent of Mg if $x = 5$. (c) Calculate the value of x if the mass percent of Mg is 16.5%.

•• **2.157** In a sample of any metal, spherical atoms pack closely together, but the space between them means that the density of the sample is less than that of the atoms themselves. Iridium (Ir) is one of the densest elements: 22.56 g/cm^3. The atomic mass of Ir is 192.22 amu, and the mass of the nucleus is 192.18 amu. Determine the density (in g/cm^3) of (a) an Ir atom and (b) an Ir nucleus. (c) How many Ir atoms placed in a row would extend 1.00 cm [radius of Ir atom = 1.36 Å; radius of Ir nucleus = 1.5 femtometers (fm); V of a sphere $= \frac{4}{3}\pi r^3$]?

•• **2.158** Which of the following steps in an overall process involve(s) a physical change and which involve(s) a chemical change?