ELEMENTS, COMPOUNDS, AND CHEMICAL REACTIONS

2

From a safe distance lightning illuminates the sky and puts on a splendid show in a springtime thunderstorm. In addition to the dazzling show, lightning has the energy to cause chemical reactions to occur between the atmosphere's oxygen and nitrogen molecules. Some of these compounds are fertilizers, essential for life below, and others are compounds that we call pollutants. In this chapter we see how to use chemical reactions to summarize the interactions of elements and compounds. *(Scott Stulberg/Corbis.)*

CHAPTER OUTLINE

THIS CHAPTER IN CONTEXT Students sometimes say that to them "chemistry is a foreign language." The statement is not far from the truth. We can consider the elements in the periodic table to be our new alphabet; the formulas for compounds are the words of chemistry; and balanced equations that show how those compounds react with each other are the sentences of this new language. Learning the language of chemistry will help you succeed because you will be able to concentrate on new concepts that depend on being fluent in our new language.

In Chapter 1 we learned about the broad scope and nature of the subject of chemistry. Importantly we learned that precise and accurate measurements and calculations are central to all sciences, especially chemistry. Now that you've learned these introductory concepts we turn our attention to atoms, elements, chemical compounds, and chemical reactions.

Our study begins with *Dalton's atomic theory*, a theory that has its roots in two basic laws of nature. The law of conservation of mass and the law of definite proportions are the foundation of one of the most important scientific theories.

The atomic theory leads us to the study of the atom's basic parts, *electrons, protons,* and *neutrons,* and how one atom and its isotopes differ from another. *Mendeleev's periodic table* on the other hand is a storehouse of relationships, trends, and similarities between the elements. When elements combine to form a compound, they do so in two broad general ways, either by the sharing of electrons between atoms to make *molecular compounds* or by the transfer of one or more electrons from one atom to another forming *ions* and *ionic compounds.* Compounds are described by *chemical formulas* and their corresponding names. How compounds and elements react with each other is described by a *balanced chemical equation.*

This chapter has three principal goals. The first is to teach you about the atomic theory so you develop an appreciation of scientific theories in general. The second is to help you understand the nature of ionic and molecular substances, how they are formed, and some of their properties. This includes understanding how to interpret chemical formulas and the basics of balancing equations. The third is to introduce you to *chemical nomenclature*—the system used to name chemical compounds. Being able to describe compounds by name is essential for communication among scientists. Therefore, we urge you to make the effort to learn how to name compounds and how to translate chemical names into chemical formulas.

As with the preceding chapter, you may already be familiar with some of the topics we discuss here. Nevertheless, be sure to study them thoroughly. By doing so you begin to build your store of factual knowledge that will enable you to more easily interpret and understand advanced topics as we get to them in later chapters.

2.1 Elements and Atoms Are Described by Dalton's Atomic Theory

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

Dalton's atomic theory explained chemical laws

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

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

Laws of chemical combination evolved from experimental observations

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

TOOLS
Law of conservation of mass and law of definite proportions

> **Law of Conservation of Mass.** No detectable gain or loss of mass occurs in chemical reactions. Mass is *conserved*.
>
> **Law of Definite Proportions.** In a given chemical compound, the elements are always combined in the same proportions by mass.

Notice that both of these laws refer to the masses of substances because the balance was one of the few chemical instruments in those times. Earlier, in our definition of matter, we noted that mass is a measure of the amount of matter in an object. Recall that mass and weight are not the same and you should be careful to use these terms correctly.

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

We actually used the **law of definite proportions** on page 6 when we defined a compound as a substance in which two or more elements are chemically combined in a *definite fixed proportion by mass*. Thus, if we decompose samples of water (a compound) into the elements oxygen and hydrogen, we always find that the ratio of oxygen to hydrogen, *by mass*, is 8 to 1. In other words, the mass of oxygen obtained is always eight times the mass of hydrogen.

9 g water → 1 g hydrogen, 8 g oxygen   18 g water → 2 g hydrogen, 16 g oxygen

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

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

100 g oxygen 1 g hydrogen	chemical reaction → occurs consuming 8 g oxygen 1 g hydrogen	9 g water 92 g oxygen
mixture before reaction		mixture after reaction

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

EXAMPLE 2.1
Applying the Law of Definite Proportions

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

ANALYSIS: As you learned in Chapter 1, much of the effort in solving a chemistry problem is devoted to determining which concepts have to be applied. We view these concepts as *tools*, each with its specific uses when applied to problem solving. Our goal in this Analysis step is to identify which tools we need and how we will apply them.

Let's begin by examining the problem and asking a question: What have we learned that relates the masses of elements in two samples of the same compound? We've described two tools relating to masses, the laws of conservation of mass and definite proportions. The law of conservation of mass concerns only the total mass of the chemicals in a reaction, so it doesn't seem to help us here. The law of definite proportions does seem to apply since it concerns the mathematical relationships of elements within a compound no matter where the sample came from. It states that the proportions of the elements by mass must be the same in both samples; so the law of definite proportions is the tool we need to apply. The law tells us that the ratio of grams of Mo to grams of S must be the same in both samples. To solve the problem, then, we will set up the mass ratios for the two samples. In the ratio for the second sample, the mass of molybdenum will be an unknown quantity. We'll equate the two ratios and solve for the unknown quantity.

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

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

In the second sample, we know the mass of S (2.50 g) and we want to find the mass of Mo (the unknown is *mass of Mo*). The mass ratio of Mo to S in the second sample is therefore

$$\frac{\textit{mass of Mo}}{2.50 \text{ g S}}$$

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

$$\frac{\textit{mass of Mo}}{2.50 \text{ g S}} = \frac{1.50 \text{ g Mo}}{1.00 \text{ g S}}$$

Solving for the *mass of Mo* gives

$$\textit{Mass of Mo} = 2.50 \cancel{\text{g S}} \times \frac{1.50 \text{ g Mo}}{1.00 \cancel{\text{g S}}} = 3.75 \text{ g Mo}$$

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

Practice Exercise 1: Cadmium sulfide is a yellow compound that is used as a pigment in artist's oil colors. A sample of this compound is composed of 1.25 g of cadmium and 0.357 g of sulfur. If a second sample of the same compound contains 3.50 g of sulfur, how many grams of cadmium does it contain? (Hint: Identify the law and write its mathematical form as it applies to this problem.)

Practice Exercise 2: Several samples of compounds containing only iron and sulfur were analyzed by taking the sample and heating it strongly to produce gaseous sulfur oxides and metallic iron. Which of the following compounds are the same and which are different?

Sample	Mass of compound before heating	Mass of iron after heating
A	25.36 g	16.11 g
B	15.42 g	8.28 g
C	7.85 g	4.22 g
D	11.87 g	7.54 g

The atomic theory was proposed by John Dalton

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

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

Dalton's Atomic Theory

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

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

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

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

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

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

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

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

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

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

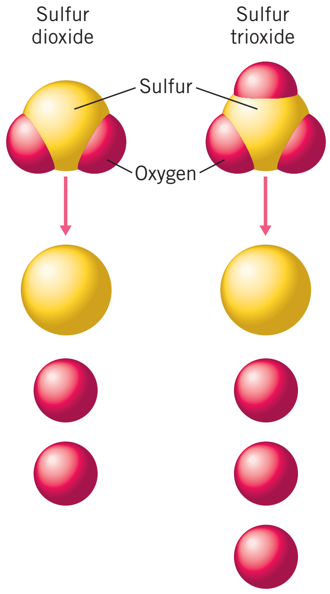


FIG. 2.1 Oxygen compounds of sulfur demonstrate the law of multiple proportions. Illustrated here are molecules of sulfur trioxide and sulfur dioxide. Each has one sulfur atom, and therefore the same mass of sulfur. The oxygen ratio is 3 to 2, both by atoms and by mass.

Dalton's theory explains the law of multiple proportions in a very simple way. Suppose a molecule of sulfur trioxide contains one sulfur and three oxygen atoms, and a molecule of sulfur dioxide contains one sulfur and two oxygen atoms (Figure 2.1). If we had just one molecule of each, then our samples each would have one sulfur atom and therefore the same mass of sulfur. Then, comparing the oxygen atoms, we find they are in a numerical ratio of 3 to 2. But because oxygen atoms all have the same mass, the mass ratio must also be 3 to 2.

The law of multiple proportions was not known before Dalton presented his theory, and its discovery demonstrates the scientific method in action. Experimental data suggested to Dalton the existence of atoms, and the atomic theory suggested the relationships that we now call the law of multiple proportions. Repeated experimental tests have uncovered no instances where the law of multiple proportions fails. These successful tests added great support to the atomic theory. In fact, for many years the law was one of the strongest arguments in favor of the existence of atoms.

Modern experimental evidence exists for atoms

Atoms are so incredibly tiny that even the most powerful optical microscopes are unable to detect them. In recent times, though, scientists have developed very sensitive instruments that are able to map the surfaces of solids with remarkable resolution. One such instrument is called a **scanning tunneling microscope.** It was invented in the early 1980s by Gerd Binnig and Heinrich Rohrer and earned them the 1986 Nobel Prize in Physics. With this instrument, the tip of a sharp metal probe is brought very close to an electrically conducting surface and an electric current bridging the gap is begun. The flow of current is extremely sensitive to the distance between the tip of the probe and the sample. As the tip is moved across the surface, the height of the tip is continually adjusted to keep the current flow constant. By accurately recording the height fluctuations of the tip, a map of the hills and valleys on the surface is obtained. The data are processed using a computer to reveal images such as that shown in Figure 2.2.

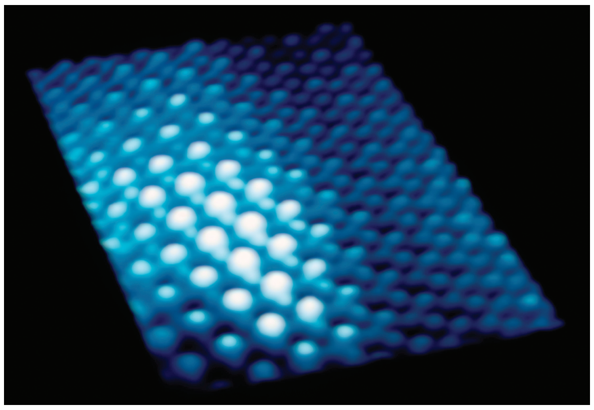


FIG. 2.2 Individual atoms can be imaged using a scanning tunneling microscope. This STM micrograph reveals the pattern of individual atoms of palladium deposited on a graphite surface. Palladium is a silvery white metal used in alloys such as white gold and dental crowns. *(Eurelios/Phototake.)*

2.2 Atoms are composed of subatomic particles

The earliest theories about atoms imagined them to be indestructible and totally unable to be broken into smaller pieces. However, as you probably know, atoms are not quite as indestructible as Dalton had thought. During the late 1800s and early 1900s, experiments were performed that demonstrated that atoms are composed of **subatomic particles.** (For some of the details about these experiments, see Facets of Chemistry 2.1.) From this work the current theoretical model of atomic structure evolved. We will examine it in general terms in this chapter. A more detailed discussion of atomic structure will follow in Chapter 7.

Protons, neutrons, and electrons are subatomic particles

Physicists have discovered a large number of subatomic particles, but protons, neutrons, and electrons are the only ones that will concern us at this time.

Protons are in all nuclei. Except for ordinary hydrogen, all nuclei also contain neutrons.

Experiments have shown that atoms are composed of three principal kinds of subatomic particles: **protons, neutrons,** and **electrons.** Experiments also revealed that at the center of an atom there exists a very tiny, extremely dense core called the **nucleus,** which is where an atom's protons and neutrons are found. Because they are found in nuclei, protons and neutrons are sometimes called **nucleons.** The electrons in an atom surround the nucleus and fill the remaining volume of the atom. (*How* the electrons are distributed around the nucleus is the subject of Chapter 7.) The properties of the subatomic particles are summarized in Table 2.1, and the general structure of the atom is illustrated in Figure 2.3.

Notice that two of the subatomic particles carry electrical charges. Protons carry a single unit of **positive charge** and electrons carry one unit of the opposite charge, a **negative charge.** Two particles that have the same electrical charge repel each other and two particles that have opposite charges will experience an attractive force. In an atom the negatively charged electrons are attracted to positively charged protons. In fact, it is this attraction that holds the electrons around the nucleus. Neutrons have no charge and are said to be electrically neutral (hence the name *neutron*).

Because of their identical charges, electrons repel each other. The repulsions between the electrons keep them spread out throughout the volume of the atom, and it is the *balance* between the attractions the electrons feel toward the nucleus and the repulsions they feel toward each other that controls the sizes of atoms.

The binding energy of nucleons is discussed in Chapter 20.

Protons also repel each other, but they are able to stay together in the small volume of the nucleus because their repulsions are apparently offset by powerful nuclear forces that involve other subatomic particles we will not study.

Matter as we generally find it in nature appears to be electrically neutral, which means that it contains equal numbers of positive and negative charges. Therefore, *in a neutral atom, the number of electrons must equal the number of protons.*

The proton and neutron are much more massive than the electron, so in any atom almost all of the atomic mass is contributed by the particles that are found in the nucleus. (The mass of an electron is only about 1/1800 of that of a proton or neutron.) It is also interesting to note, however, that the diameter of an atom is approximately 10,000 times the diameter of its nucleus, so almost all the *volume* of an atom is occupied by its electrons, which fill the space around the nucleus. (To place this on a more meaningful scale, if the nucleus were 1 ft in diameter, it would lie at the center of an atom with a diameter of approximately 1.9 miles!)

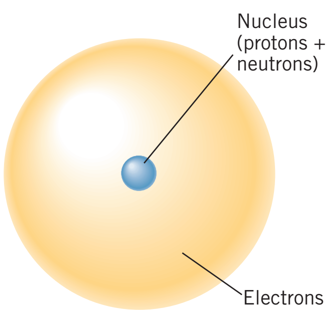


FIG. 2.3 The internal structure of an atom. An atom is composed of a tiny nucleus that holds all the protons and neutrons; the electrons fill the space outside the nucleus.

TABLE 2.1 Properties of Subatomic Particles

Particle	Mass (g)	Electrical Charge	Symbol
Electron	9.109383×10^{-28}	1−	$^{0}_{-1}e$
Proton	$1.6726217 \times 10^{-24}$	1+	$^{1}_{1}H^{+}$, $^{1}_{1}p$
Neutron	$1.6749273 \times 10^{-24}$	0	$^{1}_{0}n$

FACETS OF CHEMISTRY 2.1

Experiments Leading to the Discovery of Subatomic Particles

Our current knowledge of atomic structure was pieced together from facts obtained from experiments by scientists that began in the nineteenth century. In 1834, Michael Faraday discovered that the passage of electricity through aqueous solutions could cause chemical changes, which was the first hint that matter was electrical in nature. Later in that century, scientists began to experiment with *gas discharge tubes* in which a high-voltage electric current was passed through a gas at low pressure in a glass tube (Figure 1). Such a tube is fitted with a pair of metal *electrodes* and when the electricity begins to flow between them, the gas in the tube glows. This flow of electricity is called an *electric discharge*, which is how the tubes got their names. (Modern neon signs work this way.)

The physicists who first studied this phenomenon did not know what caused the tube to glow, but tests soon revealed that negatively charged particles were moving from the negative electrode (the *cathode*) to the positive electrode (the *anode*). According to legend, it was Benjamin Franklin who decided which electrode was positive and which was negative. The physicists called these emissions *rays,* and because the rays came from the cathode, they were called *cathode rays*.

In 1897, the British physicist J. J. Thomson constructed a special gas discharge tube to make quantitative measurements of the properties of cathode rays. In some ways, the *cathode ray tube* he used was similar to a TV picture tube, as Figure 2 shows. In Thomson's tube, a beam of cathode rays was focused on a glass surface coated with a phosphor that glows when the cathode rays strike it (point 1). The cathode ray beam passed between the poles of a magnet and between a pair of metal electrodes that could be given electrical charges. The magnetic field tends to bend the beam in one direction (to point 2) whereas the charged electrodes bend the beam in the opposite direction (to point 3). By adjusting the charge on the electrodes, the two effects can be made to cancel, and from the amount of charge on the electrodes required to balance the effect of the magnetic field, Thomson was able to calculate the first bit of quantitative information about a cathode ray particle—the ratio of its charge to its mass (often expressed as e/m, where e stands for charge and m stands for mass). The charge-to-mass ratio has a value of -1.76×10^8 coulombs/gram, where the coulomb (C) is a standard unit of electrical charge and the negative sign reflects the negative charge on the particle.

Many experiments were performed using the cathode ray tube and they demonstrated that cathode ray particles are in all matter. They are, in fact, *electrons*.

Measuring the Charge and Mass of the Electron. In 1909, a researcher at the University of Chicago, Robert Millikan, designed a clever experiment that enabled him to measure the electron's charge (Figure 3). During an experiment he would spray a fine mist of oil droplets above a pair of parallel metal plates, the top one of which had a small hole in it. As the oil drops settled, some would pass through this hole into the space between the plates, where he would irradiate them briefly with X rays. The X rays knocked electrons off molecules in the air, and the electrons became attached to the oil drops, which thereby were given an electrical charge. By observing the rate of fall of the charged drops both when the metal plates were electrically charged and when they were not, Millikan was able to calculate the amount of charge carried by each drop. When he examined his results, he found that all the values he obtained were whole-number multiples of -1.60×10^{-19} C. He reasoned that since a drop could only pick up whole numbers of electrons, this value must be the charge carried by each individual electron.

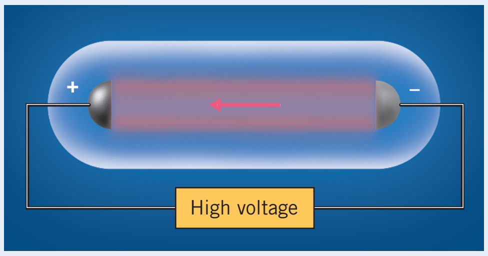


FIG. 1 **A gas discharge tube.** Cathode rays flow from the negatively charged cathode to the positively charged anode.

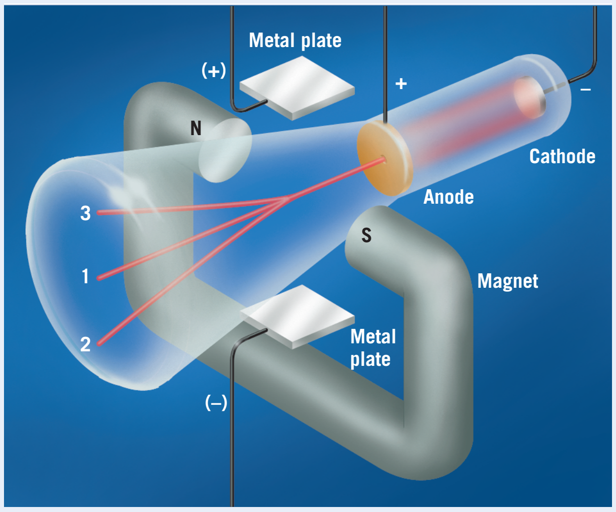


FIG. 2 **Thomson's cathode ray tube.** This device was used to measure the charge-to-mass ratio for the electron.

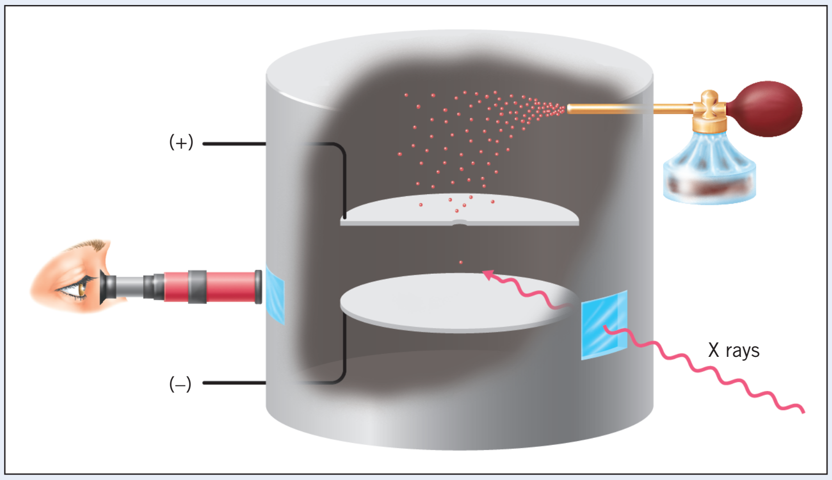


FIG. 3 Millikan's oil drop experiment. Electrons, which are ejected from air molecules by the X rays, are picked up by very small drops of oil falling through the tiny hole in the upper metal plate. By observing the rate of fall of the charged oil drops, with and without electrical charges on the metal plates, Millikan was able to calculate the charge carried by an electron.

Once Millikan had measured the electron's charge, its mass could then be calculated from Thomson's charge-to-mass ratio. This mass was calculated to be 9.09×10^{-28} g. More precise measurements have since been made, and the mass of the electron is currently reported to be 9.109383×10^{-28} g.

Discovery of the Proton. The removal of electrons from an atom gives a positively charged particle (called an *ion*). To study these, a modification was made in the construction of the cathode ray tube to produce a new device called a *mass spectrometer*. This apparatus is described in Facets of Chemistry 2.2 on page 47 and was used to measure the charge-to-mass ratios of positive ions. These ratios were found to vary, depending on the chemical nature of the gas in the discharge tube, showing that their masses also varied. The lightest positive particle observed was produced when hydrogen was in the tube, and its mass was about 1800 times as heavy as an electron. When other gases were used, their masses always seemed to be whole-number multiples of the mass observed for hydrogen atoms. This suggested the possibility that clusters of the positively charged particles made from hydrogen atoms made up the positively charged particles of other gases. The hydrogen atom, minus an electron, thus seemed to be a fundamental particle in all matter and was named the *proton*, after the Greek word *proteios,* meaning "of first importance."

Discovery of the Atomic Nucleus

Early in the twentieth century, Hans Geiger and Ernest Marsden, working under Ernest Rutherford at Great Britain's Manchester University, studied what happened when *alpha rays* hit thin metal foils. Alpha rays are composed of particles having masses four times those of the proton and bearing two positive charges; they are emitted by certain unstable atoms in a phenomenon called *radioactivity*. Most of the alpha particles sailed right on through as if the foils were virtually empty space (Figure 4). A significant number of alpha particles,

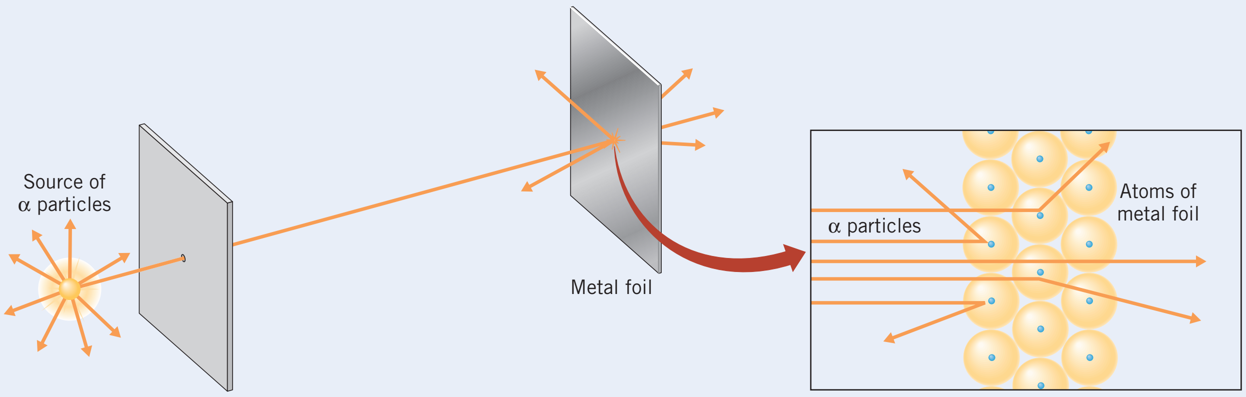


FIG. 4 Rutherford's alpha-particle experiment. Alpha particles are scattered in all directions by a thin metal foil. Some hit something very massive head-on and are deflected backward. Many sail through. Some, making near misses with the massive "cores" (nuclei), are still deflected, because alpha particles have the same kind of charge (+) as these cores.

however, were deflected at very large angles. Some were even deflected backward, as if they had hit stone walls. Rutherford was so astounded that he compared the effect to that of firing a 15 in. artillery shell at a piece of tissue paper and having it come back and hit the gunner! From studying the angles of deflection of the particles, Rutherford reasoned that only something extraordinarily massive and positively charged could cause such an occurrence. Since most of the alpha particles went straight through, he further reasoned that the metal atoms in the foils must be mostly empty space. Rutherford's ultimate conclusion was that virtually all of the mass of an atom must be concentrated in a particle having a very small volume located in the center of the atom. He called this massive particle the atom's *nucleus*.

Discovery of the Neutron

From the way alpha particles were scattered by a metal foil, Rutherford and his students were able to estimate the number of positive charges on the nucleus of an atom of the metal. This had to be equal to the number of protons in the nucleus, of course. But when they computed the nuclear mass based on this number of protons, the value always fell short of the actual mass. In fact, Rutherford found that only about half of the nuclear mass could be accounted for by protons. This led him to suggest that there were other particles in the nucleus that had a mass close to or equal to that of a proton, but with no electrical charge. This suggestion initiated a search that finally ended in 1932 with the discovery of the *neutron* by Sir James Chadwick, a British physicist.

Atomic numbers define elements and mass numbers describe isotopes

TOOLS *The number of electrons, protons, and neutrons in atoms*

What distinguishes one element from another is the number of protons in the nuclei of its atoms, because *all the atoms of a particular element have an identical number of protons.* In fact, this allows us to redefine an **element** as *a substance whose atoms all contain the identical number of protons.* Thus, each element has associated with it a unique number, which we call its **atomic number** (***Z***), that equals the number of protons in the nuclei of any of its atoms.

$$\text{Atomic number } (Z) = \text{number of protons}$$

Most elements exist in nature as mixtures of similar atoms called *isotopes* that differ only in mass. What makes isotopes of the same element different are the numbers of neutrons in their nuclei. *The* **isotopes** *of a given element have atoms with the same number of protons but different numbers of neutrons.* The numerical sum of the protons and neutrons in the atoms of a particular isotope is called the **mass number** (***A***) of the isotope.

$$\text{Isotope mass number } (A) = (\text{number of protons}) + (\text{number of neutrons})$$

Therefore, every isotope is fully defined by two numbers, its atomic number and its mass number. Sometimes these numbers are added to the left of the chemical symbol as a subscript and a superscript, respectively. Thus, if *X* stands for the chemical symbol for the element, an isotope of *X* is represented as

$$^{A}_{Z}X$$

The isotope of uranium used in nuclear reactors, for example, can be symbolized as follows:

mass number (protons + neutrons) → $^{235}_{92}U$
atomic number (number of protons) →
uranium-235

TOOLS *Atomic symbols for isotopes*

As indicated, the name of this isotope is uranium-235 or U-235. Each neutral atom contains 92 protons and $(235 - 92) = 143$ neutrons as well as 92 electrons. In writing the symbol for the isotope, the atomic number is often omitted because it is redundant. Every atom of uranium has 92 protons, and every atom that has 92 protons is an atom of uranium. Therefore, this uranium isotope can be represented simply as ^{235}U.

It is useful to remember that for a neutral atom, the atomic number equals both the number of protons and the number of electrons.

In naturally occurring uranium, a more abundant isotope is ^{238}U. Atoms of this isotope also have 92 protons, but the number of neutrons is 146. Thus, atoms of ^{235}U and ^{238}U have the identical number of protons but differ in the numbers of neutrons.

EXAMPLE 2.2
Counting Protons, Neutrons, and Electrons

How many electrons, protons, and neutrons does the Cr-52 isotope have?

ANALYSIS: This problem asks for all three of the major subatomic particles in the chromium isotope that has a nominal mass of 52. Therefore, all three of the tools for subatomic particles must be used:

$$\text{Protons} = \text{atomic number} = Z$$
$$\text{Electrons} = \text{atomic number} = Z$$
$$\text{Neutrons} = \text{mass number} - \text{atomic number} = A - Z$$

SOLUTION: Applying the tools we find that $Z = 24$ and $A = 52$ and we conclude

$$\text{Protons} = 24 \qquad \text{Electrons} = 24 \qquad \text{Neutrons} = 52 - 24 = 28$$

IS THE ANSWER REASONABLE? One check is to be sure that the sum of the number of protons and neutrons is the mass number. A second check is that the number of any of the particles is not larger than the mass (the largest number given in the problem) and in most cases the number of electrons, protons, or neutrons is usually close to half of the mass. Our answers fulfill these conditions.

Practice Exercise 3: Write the symbol for the isotope of plutonium (Pu) that contains 146 neutrons. How many electrons does it have? (Hint: Review the tools for writing isotope symbols and counting electrons.)

Practice Exercise 4: How many protons, neutrons, and electrons are in each atom of $^{35}_{17}Cl$?

Practice Exercise 5: In the previous exercise, can we discard the 35 or the 17 or both from the symbol without losing the ability to solve the problem? Explain your reasoning.

Relative atomic masses of elements can be found

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

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

Hydrogen (H) combines with the element fluorine (F) to form the compound hydrogen fluoride. Each molecule of this compound contains one atom of hydrogen and one atom of fluorine, which means that in *any* sample of this substance the fluorine-to-hydrogen *atom ratio* is always 1 to 1. It is also found that when a sample of hydrogen fluoride is decomposed, the mass of fluorine obtained is always 19.0 times larger than the mass of hydrogen, so the fluorine-to-hydrogen *mass ratio* is always 19.0 to 1.00.

F-to-H atom ratio: 1 to 1
F-to-H mass ratio: 19.0 to 1.00

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

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

TOOLS Atomic masses are relative

Carbon-12 is the standard on the atomic mass scale

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

□ The atomic mass unit is sometimes called a dalton, 1 u = 1 dalton.

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

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

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

Chemists generally work with whatever *mixture* of isotopes comes with a given element as it occurs naturally. Because the composition of this isotopic mixture is very nearly constant regardless of the source of the element, we can speak of an *average atom* of the element—average in terms of mass. For example, naturally occurring hydrogen is a mixture of two isotopes in the relative proportions given in the margin. The "average atom" of the element hydrogen, as it occurs in nature, has a mass that is 0.083992 times that of a ^{12}C atom. Since $0.083992 \times 12.000\ u = 1.0079\ u$, the average atomic mass of hydrogen is 1.0079 u. Notice that this average value is just a little larger than the atomic mass of ^{1}H because naturally occurring hydrogen also contains a little ^{2}H as shown in Table 2.2.

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

TABLE 2.2 Abundance of Hydrogen Isotopes

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

In general, the mass number of an isotope differs slightly from the atomic mass of the isotope. For instance, the isotope ^{35}Cl has an atomic mass of 34.968852 u. In fact, the *only* isotope that has an atomic mass equal to its mass number is ^{12}C; *by definition* the mass of this atom is exactly 12 u.

Average atomic masses can be calculated from isotopic abundances

Originally, the relative atomic masses of the elements were determined in a way similar to that described for hydrogen and fluorine in our earlier discussion. A sample of a compound was analyzed and from the formula of the substance the relative atomic masses were calculated. These were then adjusted to place them on the unified atomic mass scale. In modern times, methods have been developed to measure very precisely both the relative abundances of the isotopes of the elements and their atomic masses. (See Facets of Chemistry 2.2.) This kind of information has permitted the calculation of more precise values of the average atomic masses, which are found in the table on the inside front cover of the book. Example 2.3 illustrates how this calculation is done.

EXAMPLE 2.3

Calculating Average Atomic Masses from Isotopic Abundances

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

ANALYSIS: In any natural sample containing many atoms of chlorine, 75.77% of the mass is contributed by atoms of ^{35}Cl and 24.23% is contributed by atoms of ^{37}Cl. This means that when we calculate the mass of the "average atom," we have to proportion it according to both the masses of the isotopes and their relative abundances. It is convenient to imagine an "average atom" to be composed of 75.77% of ^{35}Cl and 24.23% of ^{37}Cl. (Keep in mind, of course, that such an atom doesn't really exist.) We also recall that when we need to use percentages in calculations, we must divide the percentage by 100 to obtain a decimal number. This decimal number, when multiplied by the isotope mass, will tell us how much of the average mass is contributed by that isotope. All we need to do is add the contributions for all isotopes to obtain the average mass.

SOLUTION: We will calculate 75.77% of the mass of an atom of ^{35}Cl, which is the contribution of this isotope to the "average atom":

$$\text{contribution of } ^{35}\text{Cl} = \frac{75.77\%\ ^{35}\text{Cl} \times 34.9689\ \text{u}}{100\ \%} = 26.496\ \text{u}$$

and for the ^{37}Cl, its contribution is

$$\text{contribution of } ^{37}\text{Cl} = \frac{24.23\%\ ^{37}\text{Cl} \times 36.9659\ \text{u}}{100\%} = 8.9568\ \text{u}$$

Then we add these contributions to give us the total mass of the "average atom."

$$26.496\ \text{u} + 8.957\ \text{u} = 35.453\ \text{u rounded to } 35.45\ \text{u}$$

Notice that in a two-step problem we kept one extra significant figure until the final rounding to four significant figures.

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

Practice Exercise 6: Aluminum atoms have a mass that is 2.24845 times that of an atom of ^{12}C. What is the atomic mass of aluminum? (Hint: Recall that we have a tool that gives the relationship between the atomic mass unit and ^{12}C.)

Practice Exercise 7: How much heavier than an atom of ^{12}C is the average atom of naturally occurring copper? Refer to the table inside the front cover of the book for the necessary data.

Practice Exercise 8: Naturally occurring boron is composed of 19.8% of ^{10}B and 80.2% of ^{11}B. Atoms of ^{10}B have a mass of 10.0129 u and those of ^{11}B have a mass of 11.0093 u. Calculate the average atomic mass of boron.

FACETS OF CHEMISTRY 2.2

The Mass Spectrometer and the Experimental Measurement of Atomic Masses

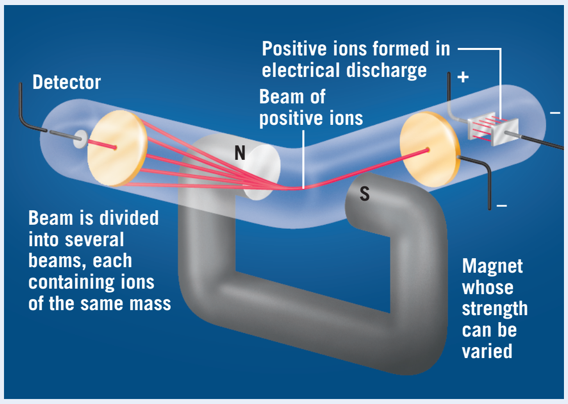
When a spark is passed through a gas, electrons are knocked off the gas molecules. Because electrons are negatively charged, the particles left behind carry positive charges; they are called *positive ions*. These positive ions have different masses, depending on the masses of the molecules from which they are formed. Thus, some molecules have large masses and give heavy ions, while others have small masses and give light ions.

The device that is used to study the positive ions produced from gas molecules is called a *mass spectrometer* (illustrated in the figure at the right). In a mass spectrometer, positive ions are created by passing an electrical spark (called an *electric discharge*) through a sample of the particular gas being studied. As the positive ions are formed, they are attracted to a negatively charged metal plate that has a small hole in its center. Some of the positive ions pass through this hole and travel onward through a tube that passes between the poles of a powerful magnet.

One of the properties of charged particles, both positive and negative, is that their paths become curved as they pass through a magnetic field. This is exactly what happens to the positive ions in the mass spectrometer as they pass between the poles of the magnet. However, the extent to which their paths are bent depends on the masses of the ions. This is because the path of a heavy ion, like that of a speeding cement truck, is difficult to change, but the path of a light ion, like that of a motorcycle, is influenced more easily. As a result, heavy ions emerge from between the magnet's poles along different lines than the lighter ions. In effect, an entering beam containing ions of different masses is sorted by the magnet into a number of beams, each containing ions of the same mass. This spreading out of the ion beam thus produces an array of different beams called a *mass spectrum*.

In practice, the strength of the magnetic field is gradually changed, which sweeps the beams of ions across a detector located at the end of the tube. As a beam of ions strikes the detector, its intensity is measured and the masses of the particles in the beam are computed based on the strength of the magnetic field and the geometry of the apparatus.

Among the benefits derived from measurements using the mass spectrometer are very accurate isotopic masses and relative isotopic abundances. These serve as the basis for the very precise values of the atomic masses that you find in the periodic table.



2.3 The Periodic Table Is Used to Organize and Correlate Facts

When we study different kinds of substances, we find that some are elements and others are compounds. Among compounds, some are composed of discrete molecules. Others, as you will learn, are made up of atoms that have acquired electrical charges. For elements such as sodium and iron we mentioned their metallic properties. On the other hand, chlorine and oxygen are not classified as metals. If we were to continue on this way, without attempting to build our subject around some central organizing structure, it would not be long before we became buried beneath a mountain of information in the form of seemingly unconnected facts.

The need for organization was recognized by many early chemists, and there were numerous attempts to discover relationships among the chemical and physical properties of the elements. A number of different sequences of elements were tried in the search for some sort of order or pattern. A few of these arrangements came quite close, at least in some respects, to our current periodic table, but either they were flawed in some way or they were presented to the scientific community in a manner that did not lead to their acceptance.

Mendeleev created the first periodic table

The periodic table we use today is based primarily on the efforts of a Russian chemist, Dmitri Ivanovich Mendeleev (1834–1907), and a German physicist, Julius Lothar Meyer (1830–1895). Working independently, these scientists developed similar periodic tables only a few months apart in 1869. Mendeleev is usually given the credit, however, because he had the good fortune to publish first.

Mendeleev was preparing a chemistry textbook for his students at the University of St. Petersburg. Looking for some pattern among the properties of the elements, he found that when he arranged them in order of increasing atomic mass, similar chemical properties were repeated over and over again at regular intervals. For instance, the elements lithium (Li), sodium (Na), potassium (K), rubidium (Rb), and cesium (Cs) are soft metals that are very reactive toward water. They form compounds with chlorine that have a 1-to-1 ratio of metal to chlorine. Similarly, the elements that immediately follow each of these also constitute a set with similar chemical properties. Thus, beryllium (Be) follows lithium, magnesium (Mg) follows sodium, calcium (Ca) follows potassium, strontium (Sr) follows rubidium, and barium (Ba) follows cesium. All of these elements form a water-soluble chlorine compound with a 1-to-2 metal to chlorine atom ratio. Mendeleev used such observations to construct his **periodic table,** which is illustrated in Figure 2.4.

Periodic refers to the recurrence of properties at regular intervals.

The elements in Mendeleev's table are arranged in order of increasing atomic mass. When the sequence is broken at the right places and stacked, the elements fall naturally into columns, called *groups*, in which the elements of a given group have similar chemical properties. The rows themselves are called *periods.*

Mendeleev's genius rested on his placing elements with similar properties in the same group even when this left occasional gaps in the table. For example, he placed arsenic (As) in Group V under phosphorus because its chemical properties were similar to those of phosphorus, even though this left gaps in Groups III and IV. Mendeleev reasoned, correctly, that the elements that belonged in these gaps had simply not yet been discovered. In fact, on the basis of the location of these gaps Mendeleev was able to predict with remarkable accuracy the properties of these yet-to-be-found substances. His predictions helped serve as a guide in the search for the missing elements.

FIG. 2.4 The first periodic table. Mendeleev's periodic table roughly as it appeared in 1871. The numbers next to the symbols are atomic masses.

Periods	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
1	H 1							
2	Li 7	Be 9.4	B 11	C 12	N 14	O 16	F 19	
3	Na 23	Mg 24	Al 27.3	Si 28	P 31	S 32	Cl 35.5	
4	K 39	Ca 40	— 44	Ti 48	V 51	Cr 52	Mn 55	Fe 56, Co 59 Ni 59, Cu 63
5	(Cu 63)	Zn 65	— 68	— 72	As 75	Se 78	Br 80	
6	Rb 85	Sr 87	?Yt 88	Zr 90	Nb 94	Mo 96	— 100	Ru 104, Rh 104 Pd 105, Ag 108
7	(Ag 108)	Cd 112	In 113	Sn 118	Sb 122	Te 128	I 127	
8	Cs 133	Ba 137	?Di 138	?Ce 140	—	—	—	— —
9	—	—	—	—	—	—	—	
10	—	—	?Er 178	?La 180	Ta 182	W 184	—	Os 195, Ir 197 Pt 198, Au 199
11	(Au 199)	Hg 200	Tl 204	Pb 207	Bi 208	—		
12	—	—	—	Th 231	—	U 240	—	— — — —

The elements tellurium (Te) and iodine (I) caused Mendeleev some problems. According to the best estimates at that time, the atomic mass of tellurium was greater than that of iodine. Yet if these elements were placed in the table according to their atomic masses, they would not fall into the proper groups required by their properties. Therefore, Mendeleev switched their order and in so doing violated his ordering sequence. (Actually, he believed that the atomic mass of tellurium had been incorrectly measured, but this wasn't so.)

The table that Mendeleev developed is in many ways similar to the one we use today. One of the main differences, though, is that Mendeleev's table lacks the column containing the elements helium (He) through radon (Rn). In Mendeleev's time, none of these elements had yet been found because they are relatively rare and because they have virtually no tendency to undergo chemical reactions. When these elements were finally discovered, beginning in 1894, another problem arose. Two more elements, argon (Ar) and potassium (K), did not fall into the groups required by their properties if they were placed in the table in the order required by their atomic masses. Another switch was necessary and another exception had been found. It became apparent that atomic mass was not the true basis for the periodic repetition of the properties of the elements. To determine what the true basis was, however, scientists had to await the discoveries of the atomic nucleus, the proton, and atomic numbers.

The modern periodic table arranges elements by atomic number

When atomic numbers were discovered, it was soon realized that the elements in Mendeleev's table are arranged in precisely the order of increasing atomic number. In other words, if we take atomic numbers as the basis for arranging the elements in sequence, no annoying switches are required and the elements Te and I or Ar and K are no longer a problem. The fact that it is the atomic number—the number of protons in the nucleus of an atom—that determines the order of elements in the table is very significant. We will see later that this has important implications with regard to the relationship between the number of electrons in an atom and the atom's chemical properties.

The modern periodic table is shown in Figure 2.5 and also appears on the inside front cover of the book. We will refer to the table frequently, so it is important for you to become familiar with it and with some of the terminology applied to it.

Special terminology is associated with the periodic table

TOOLS *Rows are periods*

As in Mendeleev's table, the elements are arranged in rows that we call **periods**, but here they are arranged in order of increasing atomic number. For identification purposes the periods are numbered. We will find these numbers useful later on. Below the main body of the table are two long rows of 14 elements each. These actually belong in the main body of the table following La ($Z = 57$) and Ac ($Z = 89$), as shown in Figure 2.6. They are almost always placed below the table simply to conserve space. If the fully spread-out table is printed on one page, the type is so small that it's difficult to read. Notice that in the fully extended form of the table, with all the elements arranged in their proper locations, there is a great deal of empty space. An important requirement of a detailed atomic theory, which we will get to in Chapter 7, is that it must explain not only the repetition of properties, but also why there is so much empty space in the table.

Recall that the symbol Z stands for atomic number.

TOOLS *Columns are groups or families*

Again, as in Mendeleev's table, the vertical columns are called **groups.** However, there is not uniform agreement among chemists on how they should be numbered. In an attempt to standardize the table, the International Union of Pure and Applied Chemistry (the IUPAC), an international body of scientists responsible for setting standards in chemistry, officially adopted a system in which the groups are simply numbered sequentially, 1 through 18, from left to right using Arabic numerals. Chemists in North America favor the system where the longer groups are labeled IA to VIIIA and the shorter groups are labeled IB to VIIIB in the sequence depicted in Figure 2.5. Note that Group VIIIB encompasses three columns; moreover, the sequence of the B-group elements is unique and will make sense when we learn more about the structure of the atom in Chapter 7. European chemists favor a third numbering system with roman numerals and the designation of A and B groups but with a different sequence from the North American table. In Figure 2.5 and on the

Alkali metals (except H)

Alkaline earth metals

Atomic number

1 H 1.00794

Atomic mass

Group designation

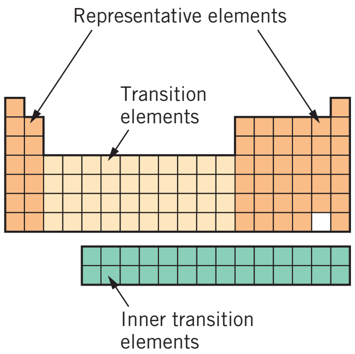
Halogens

Noble gases

Periods	IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIIIB (8)	VIIIB (9)	VIIIB (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	VIIIA (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 32.06	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	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 72.64	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 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	57 La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po [209]	85 At [210]	86 Rn [222]
7	87 Fr [223]	88 Ra [226]	89 Ac [227]	104 Rf [267]	105 Db [268]	106 Sg [271]	107 Bh [272]	108 Hs [270]	109 Mt [276]	110 Ds [281]	111 Rg [280]	112 Uub [285]	113 Uut [284]	114 Uuq [289]	115 Uup [288]	116 Uuh [293]		118 Uuo [294]

Lanthanides	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
Actinides	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np [237]	94 Pu [244]	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 [262]

FIG. 2.5 **The modern periodic table.** At room temperature, mercury and bromine are liquids. Eleven elements are gases including the noble gases and the diatomic gases of hydrogen, oxygen, nitrogen, fluorine, and chlorine. The remaining elements are solids.



inside front cover of the book, we have used both the North American labels as well as those preferred by the IUPAC. Because of the lack of uniform agreement among chemists on how the groups should be specified, we will use the North American A-group/B-group designations in Figure 2.5 when we wish to specify a particular group.

As we have already noted, the elements in a given group bear similarities to each other. Because of such similarities, groups are sometimes referred to as **families of elements.** The elements in the longer columns (the A groups) are known as the **representative elements** or **main group elements.** Those that fall into the B groups in the center of the table are called **transition elements.** The elements in the two long rows below the main body of the table are the **inner transition elements,** and each row is named after the element that it follows in the main body of the table. Thus, elements 58–71 are called the **lanthanide elements** because they follow lanthanum ($Z = 57$), and elements 90–103 are called the **actinide elements** because they follow actinium ($Z = 89$). The lanthanides are also called **rare earth metals.**

Some of the groups have acquired common names. For example, except for hydrogen, the Group IA elements are metals. They form compounds with oxygen that dissolve in water to give solutions that are strongly alkaline, or caustic. As a result, they are called the **alkali metals** or simply the *alkalis*. The Group IIA elements are also metals. Their oxygen compounds are alkaline, too, but many compounds of the Group IIA elements are unable to dissolve in water and are found in deposits in the ground. Because of their properties and where they occur in nature, the Group IIA elements became known as the **alkaline earth metals.**

On the right side of the table, in Group VIIIA, are the **noble gases.** They used to be called the **inert gases** until it was discovered that the heavier members of the group show a small degree of chemical reactivity. The term *noble* is used when we wish to suggest a very limited degree of chemical reactivity. Gold, for instance, is often referred to as a noble metal because so few chemicals are capable of reacting with it.

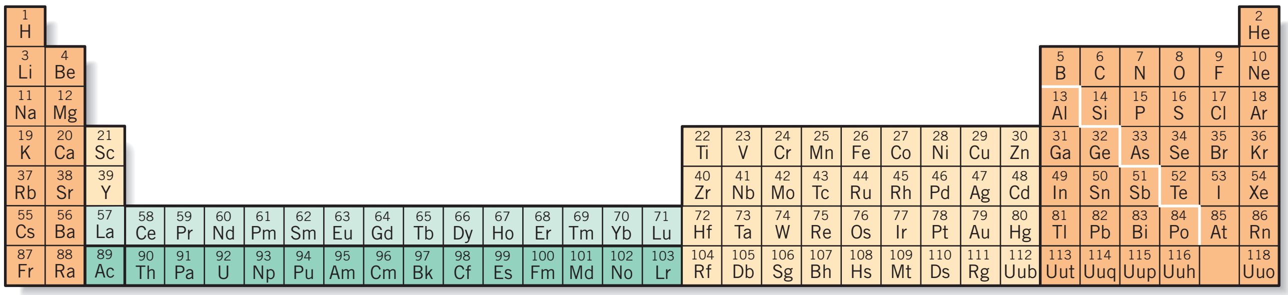


FIG. 2.6 Extended form of the periodic table. The two long rows of elements below the main body of the table in Figure 2.5 are placed in their proper places in this table.

Finally, the elements of Group VIIA are called the **halogens,** derived from the Greek word meaning "sea" or "salt." Chlorine (Cl), for example, is found in familiar table salt, a compound that accounts in large measure for the salty taste of seawater. The other groups of the representative elements have less frequently used names and we will name those groups based on the first element in the family; for example, Group VA is the **nitrogen family** and Group VIA is the **oxygen family.**

2.4 Elements Can Be Metals, Nonmetals, or Metalloids

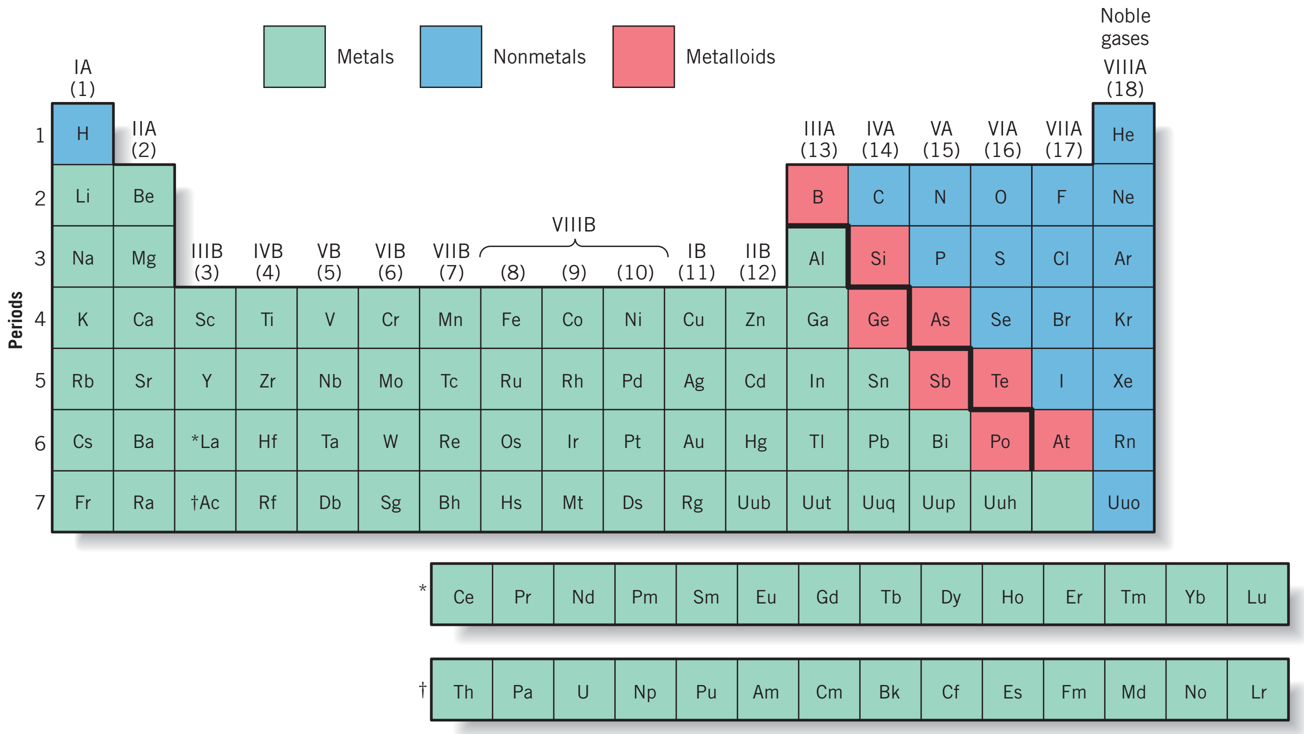
TOOLS *Periodic table metals, nonmetals, and metalloids*

The periodic table organizes all sorts of chemical and physical information about the elements and their compounds. It allows us to study systematically the way properties vary with an element's position within the table and, in turn, makes the similarities and differences among the elements easier to understand and remember.

Even a casual inspection of samples of the elements reveals that some are familiar metals and that others, equally well known, are not metals. Most of us recognize metals such as lead, iron, or gold and nonmetals such as oxygen or nitrogen. A closer look at the nonmetallic elements, though, reveals that some of them, silicon and arsenic to name two, have properties that lie between those of true metals and true nonmetals. These elements are called **metalloids.** The elements are not evenly divided into the categories of metals, nonmetals, and metalloids. (See Figure 2.7.) Most elements are metals, slightly over a dozen are nonmetals, and only a handful are metalloids.

Notice that the metalloids are grouped around the bold stair-step line that is drawn diagonally from boron (B) to astatine (At).

FIG. 2.7 Distribution of metals, nonmetals, and metalloids among the elements in the periodic table.



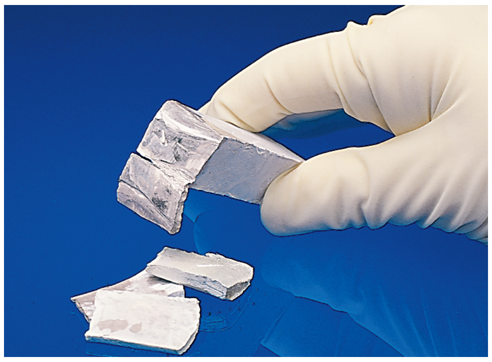


FIG. 2.8 Sodium is a metal. The freshly exposed surface of a bar of sodium reveals its shiny metallic luster. The metal reacts quickly with moisture and oxygen to form a white coating. Its high reactivity makes it dangerous to touch with bare skin. *(Michael Watson.)*

Metals have distinctive physical properties

You probably know a **metal** when you see one. Metals tend to have a shine so unique that it's called a *metallic luster*. For example, the silvery sheen of the freshly exposed surface of sodium in Figure 2.8 would most likely lead you to identify sodium as a metal even if you had never seen or heard of it before. We also know that metals conduct electricity. Few of us would hold an iron nail in our hand and poke it into an electrical outlet. In addition, we know that metals conduct heat very well. On a cool day, metals always feel colder to the touch than do neighboring nonmetallic objects because metals conduct heat away from your hand very rapidly. Nonmetals seem less cold because they can't conduct heat away as quickly and therefore their surfaces warm up faster.

Other properties that metals possess, to varying degrees, are **malleability**—the ability to be hammered or rolled into thin sheets—and **ductility**—the ability to be drawn into wire. The ability of a blacksmith to fashion horseshoes from a bar of iron (Figure 2.9) depends on the malleability of iron and steel, and the manufacture of electrical wire is based on the ductility of copper.

Thin lead sheets are used for sound-deadening because the easily deformed lead absorbs the sound vibrations.

Hardness is another physical property that we usually think of for metals. Some, such as chromium or iron, are indeed quite hard; but others, like copper and lead, are rather soft. The alkali metals such as sodium (Figure 2.8) are so soft they can be cut with a knife, but they are also so chemically reactive that we rarely get to see them as free elements.

All the metallic elements, except mercury, are solids at room temperature (Figure 2.10). Mercury's low freezing point (−39 °C) and fairly high boiling point (357 °C) make it useful as a fluid in thermometers. Most of the other metals have much higher melting points, and some are used primarily because of this. Tungsten, for example, has the highest melting point of any metal (3400 °C, or 6150 °F), which explains its use as filaments that glow white-hot in electric lightbulbs.

The chemical properties of metals vary tremendously. Some, such as gold and platinum, are very unreactive toward almost all chemical agents. This property, plus their natural beauty and rarity, makes them highly prized for use in jewelry. Other metals, however, are so reactive that few people except chemists and chemistry students ever get to see them in their "free" states. For instance, the metal sodium reacts very quickly with oxygen or moisture in the air, and its bright metallic surface tarnishes almost immediately. In contrast, compounds of sodium, such as table salt and baking soda, are quite stable and very common.

We use the term *free element* to mean an element that is not chemically combined with any other element.

Nonmetals lack the properties of metals

Substances such as plastics, wood, and glass that lack the properties of metals are said to be *nonmetallic*, and an element that has nonmetallic properties is called a **nonmetal** or **nonmetallic element.** Most often, we encounter the nonmetals in the form of compounds or mixtures of compounds. There are some nonmetals, however, that are very important to us in their elemental forms. The air we breathe, for instance, contains mostly nitrogen and oxygen. Both are gaseous, colorless, and odorless nonmetals. Since we can't see, taste, or smell them, however, it's difficult to experience their existence. (Although if you step into an atmosphere without oxygen, your body will soon tell you that something is missing!) Probably the most commonly *observed* nonmetallic element is carbon. We find it as the graphite in pencils, as coal, and as the charcoal used for barbecues. It also occurs in a more valuable form as diamond (Figure 2.11). Although diamond and graphite differ in appearance, each is a form of elemental carbon.

FIG. 2.9 Malleability of iron. A blacksmith uses the malleability of hot iron to fashion horseshoes from an iron bar. *(Stone/Getty Images.)*

Many of the nonmetals are solids at room temperature and atmospheric pressure, while many others are gases. Photographs of some of the nonmetallic elements appear in Figure 2.12. Their properties are almost completely opposite those of metals. Each of these elements lacks the characteristic appearance of a metal. They are poor conductors of heat and, with the exception of the graphite form of carbon, are also poor conductors of electricity. The electrical conductivity of graphite appears to be an accident of molecular structure, since the structures of metals and graphite are completely different.

The nonmetallic elements lack the malleability and ductility of metals. A lump of sulfur crumbles when hammered and breaks apart when pulled on. Diamond cutters rely on

FIG. 2.10 **Mercury droplet.** The metal mercury (once known as quicksilver) is a liquid at room temperature, unlike other metals, which are solids. *(OPC, Inc.)*

FIG. 2.11 **Diamond.** This gem is simply another form of the element carbon. *(Charles D. Winters/ Photo Researchers, Inc.)*

FIG. 2.12 **Some nonmetallic elements.** In the bottle on the left is dark-red liquid bromine, which vaporizes easily to give a deeply colored orange vapor. Pale green chlorine fills the round flask in the center. Solid iodine lines the bottom of the flask on the right and gives off a violet vapor. Powdered red phosphorus occupies the dish in front of the flask of chlorine, and black powdered graphite is in the watch glass. Also shown are lumps of yellow sulfur. *(Michael Watson.)*

the brittle nature of carbon when they split a gem-quality stone by carefully striking a quick blow with a sharp blade.

As with metals, nonmetals exhibit a broad range of chemical reactivity. Fluorine, for instance, is extremely reactive. It reacts readily with almost all the other elements. At the other extreme is helium, the gas used to inflate children's balloons and the blimps seen at major sporting events. This element does not react with anything, a fact that chemists find useful when they want to provide a totally *inert* (unreactive) atmosphere inside some apparatus.

Metalloids have physical properties between metals and nonmetals

The properties of metalloids lie between those of metals and nonmetals. This shouldn't surprise us since the metalloids are located between the metals and the nonmetals in the periodic table. In most respects, metalloids behave as nonmetals, both chemically and physically. However, in their most important physical property, electrical conductivity, they somewhat resemble metals. Metalloids tend to be **semiconductors;** they conduct electricity, but not nearly so well as metals. This property, particularly as found in silicon and germanium, is responsible for the remarkable progress made during the last five decades in the field of solid-state electronics. The operation of every computer, audio system, TV receiver, DVD or CD player, and AM–FM radio relies on transistors made from semiconductors. Perhaps the most amazing advance of all has been the fantastic reduction in the size of electronic components that semiconductors have allowed (Figure 2.13). To it, we owe the development of small and versatile cell phones, cameras, MP3 players, handheld calculators,

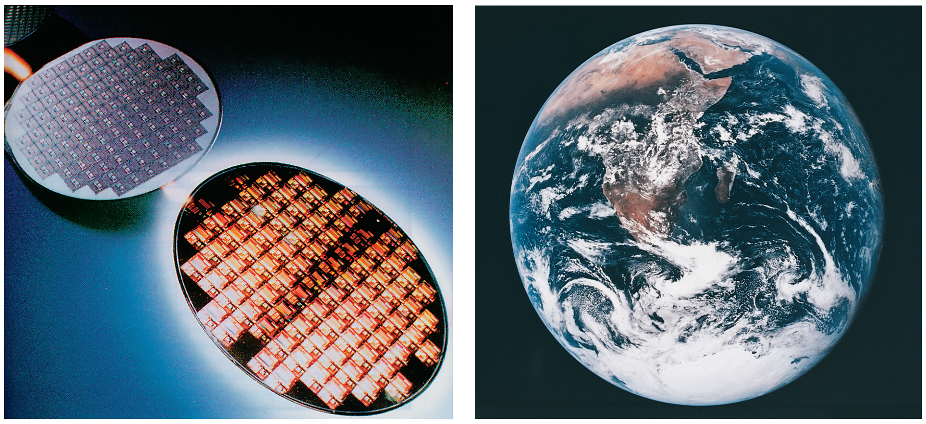


FIG. 2.13 **Modern electronic circuits rely on the semiconductor properties of silicon.** The silicon wafer shown here contains more electronic components (10 billion) than there are people on our entire planet (about 6.5 billion)! *(Courtesy Sematech; Courtesy NASA)*

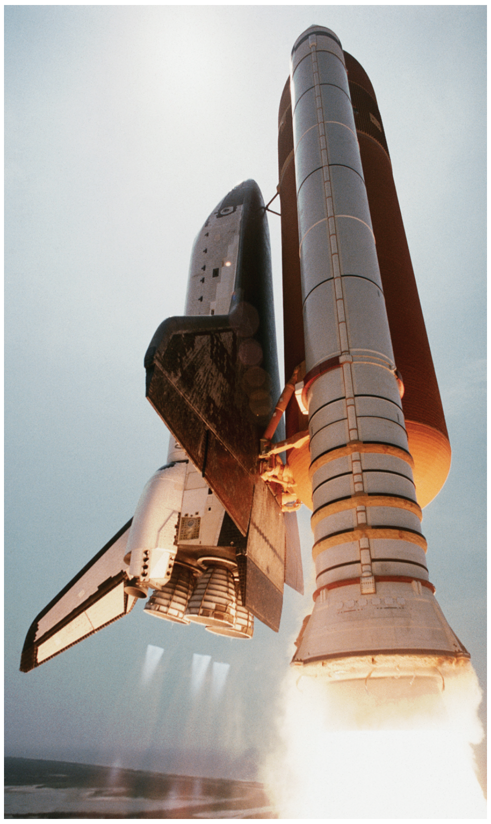


FIG. 2.14 Hydrogen reacts with oxygen to form water. The reaction of hydrogen with oxygen provides the thrust for the main rocket engines of the space shuttle shown here just after liftoff. *(Corbis Images.)*

and microcomputers. The heart of these devices is an integrated circuit that begins as a wafer of extremely pure silicon (or germanium) that is etched and chemically modified into specialized arrays of thousands of transistors.

Metallic and nonmetallic character is related to an element's position in the periodic table

The occurrence of the metalloids between the metals and the nonmetals is our first example of trends in properties within the periodic table. We will frequently see that as we move from position to position across a period or down a group, chemical and physical properties change in a more or less regular way. There are few abrupt changes in the characteristics of the elements as we scan across a period or down a group. The location of the metalloids can be seen, then, as an example of the gradual transition between metallic and nonmetallic properties. From left to right across Period 3, we go from aluminum, an element that has every appearance of a metal; to silicon, a semiconductor; to phosphorus, an element with clearly nonmetallic properties. A similar gradual change is seen going down Group IVA. Carbon is certainly a nonmetal, silicon and germanium are metalloids, and tin and lead are metals. Trends such as these are useful to spot because they help us remember properties.

2.5 Formulas and Equations Describe Substances and Their Reactions

A property possessed by nearly every element is the ability to combine with other elements to form compounds, although not all combinations appear to be possible. For example, iron reacts with oxygen to form a compound that we commonly call rust, but no compound is formed between sodium and iron.

In Chapter 1 we noted that during chemical reactions, the properties of the substances change, often dramatically, when a reaction takes place. This is certainly true in the reactions of elements to form compounds. An example is the reaction between hydrogen and oxygen to form ordinary water.

At room temperature both hydrogen and oxygen are clear colorless gases. When they are mixed and ignited, these elements combine explosively to form the familiar compound water, which of course is a liquid at room temperature. As with nearly all chemical reactions, the properties of the substances present prior to the reaction differ quite a lot from the properties of those present afterwards.

The reaction between hydrogen and oxygen is not one we would expect to encounter in our daily lives, but it does have applications. When hydrogen and oxygen are cooled to sufficiently low temperatures, they condense to form liquids that serve as the fuel for the main rocket engines of the space shuttle (Figure 2.14). Hydrogen has also been used to power nonpolluting vehicles in which its reaction with oxygen (from air) yields an exhaust containing only water vapor.

A chemical formula describes the composition of a substance

TOOLS
Chemical formulas

To describe chemical substances, both elements and compounds, we commonly use **chemical formulas,** in which chemical symbols are used to represent atoms of the elements that are present. For a **free element** (*one that is not combined with another element in a compound*) we often simply use the chemical symbol. Thus, the element sodium is represented by its symbol, Na, which is interpreted to mean one atom of sodium.

Except for the noble gases, all the free nonmetallic elements exist as molecules that contain two or more atoms. Many of those we encounter frequently occur as **diatomic molecules** (molecules composed of two atoms each). Among them are the gases hydrogen, oxygen, and nitrogen and the halogens (fluorine, chlorine, bromine, and iodine). We represent these elements with chemical formulas in which subscripts indicate the number of atoms in a molecule. Thus, the **formula** for molecular hydrogen is H_2, and those for oxygen, nitrogen, and chlorine are O_2, N_2, and Cl_2, respectively. (See Figure 2.15.) The elements that occur as diatomic molecules are listed in Table 2.3. This would be a good time to learn them because

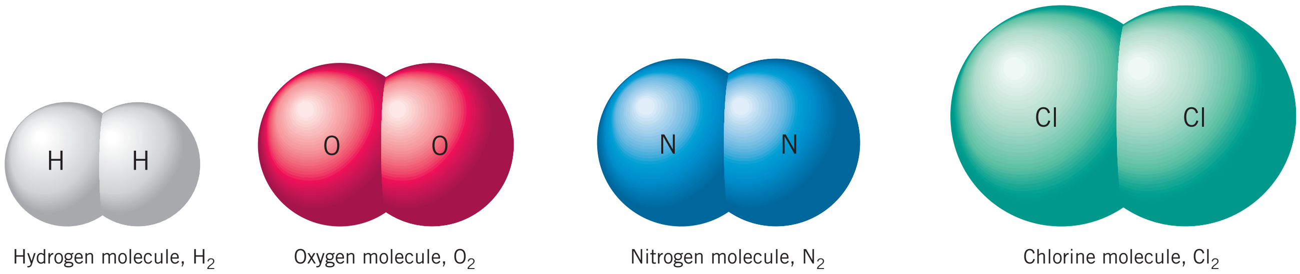


FIG. 2.15 Models that depict the diatomic molecules of hydrogen, oxygen, nitrogen, and chlorine. Each contains two atoms per molecule; their different sizes reflect differences in the sizes of the atoms that make up the molecules. The atoms are shaded by color to indicate the element (hydrogen, white; oxygen, red; nitrogen, blue; and chlorine, green).

TABLE 2.3 Elements That Occur Naturally as Diatomic Molecules

Hydrogen	H_2	Fluorine	F_2
Nitrogen	N_2	Chlorine	Cl_2
Oxygen	O_2	Bromine	Br_2
		Iodine	I_2

you will come upon them often throughout the course. Other nonmetals have their atoms arranged in even more complex combinations. Elemental sulfur, for example, contains molecules of S_8 and one form of phosphorus has molecules of P_4.

Just as chemical symbols can be used as shorthand notations for the names of elements, a chemical formula is a shorthand way of writing the name for a compound. However, *the most important characteristic of a compound's formula is that it specifies the composition of the substance.*

In the formula of a compound, each element present is identified by its chemical symbol. Table salt, for example, has the chemical formula NaCl which indicates it is composed of the elements sodium (Na) and chlorine (Cl). When more than one atom of an element is present, the number of atoms is given by a **subscript** after the symbol. For instance, the iron oxide in rust has the formula Fe_2O_3, which tells us that the compound is composed of iron (Fe) and oxygen (O), and that in this compound there are two atoms of iron for every three atoms of oxygen. When no subscript is written, we assume it to be 1, so in NaCl we find one atom of sodium (Na) for each atom of chlorine (Cl). Similarly, the formula H_2O tells us that in water there are two hydrogen atoms for every one oxygen atom, and the formula for chloroform, $CHCl_3$, indicates that one atom of carbon, one atom of hydrogen, and three atoms of chlorine have combined (Figure 2.16).

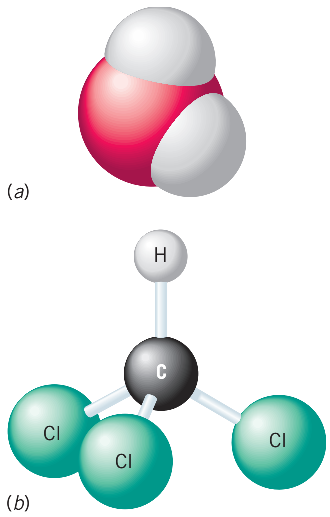
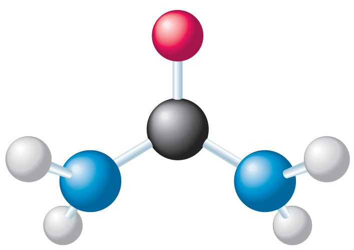


FIG. 2.16 Molecules of water and chloroform. (*a*) A space-filling model of H_2O. (*b*) A ball-and-stick model of $CHCl_3$. (Hydrogen, white; oxygen, red; carbon, black; and chlorine, green.)

For more complicated compounds, we sometimes find formulas that contain parentheses. An example is the formula for urea, $CO(NH_2)_2$, which tells us that the group of atoms within the parentheses, NH_2, occurs twice. (The formula for urea also could be written as CON_2H_4, but there are good reasons for writing certain formulas with parentheses, as you will see later.)



Ball-and-stick model of the urea molecule, $CO(NH_2)_2$.

Hydrates are crystals that contain water in fixed proportions

Certain compounds form crystals that contain water molecules. An example is ordinary plaster—the material often used to coat the interior walls of buildings. Plaster consists of crystals of calcium sulfate, $CaSO_4$, that contain two molecules of water for each $CaSO_4$. These water molecules are not held very tightly and can be driven off by heating the crystals. The dried crystals absorb water again if exposed to moisture, and the amount of water absorbed always gives crystals in which the H_2O-to-$CaSO_4$ ratio is 2 to 1. Compounds whose crystals contain water molecules in fixed ratios are quite common and are called **hydrates**. The formula for this hydrate of calcium sulfate is written $CaSO_4 \cdot 2H_2O$ to show that there are two molecules of water per $CaSO_4$. The raised dot is used to indicate that the water molecules are not bound too tightly in the crystal and can be removed.

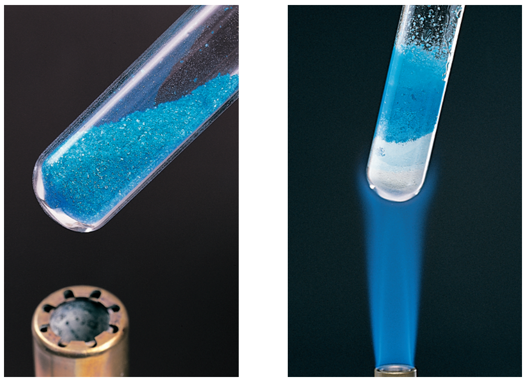


FIG. 2.17 Water can be driven from hydrates by heating. (*a*) Blue crystals of copper sulfate pentahydrate, $CuSO_4{\cdot}5H_2O$, about to be heated. (*b*) The hydrate readily loses water when heated. The light colored solid observed in the lower half of the test tube is pure $CuSO_4$. *(Richard Megna/Fundamental Photographs; Michael Watson)*

Sometimes the *dehydration* (removal of water) of hydrate crystals produces changes in color. An example is copper sulfate, which is sometimes used as an agricultural fungicide. Copper sulfate forms blue crystals with the formula $CuSO_4{\cdot}5H_2O$ in which there are five water molecules for each $CuSO_4$. When these blue crystals are heated, most of the water is driven off and the solid that remains, now nearly pure $CuSO_4$, is almost white (Figure 2.17). If left exposed to the air, the $CuSO_4$ will absorb moisture and form blue $CuSO_4{\cdot}5H_2O$ again.

When all the water is removed, the solid is said to be **anhydrous**, meaning "without water."

Counting atoms in formulas is a necessary skill

Counting the number of atoms of the elements in a chemical formula is an operation you will have to perform many times, so let's look at an example.

EXAMPLE 2.4
Counting Atoms in Formulas

How many atoms of each element are represented by the formulas (a) $Al_2(SO_4)_3$ and (b) $CoCl_2{\cdot}6H_2O$?

ANALYSIS: The essential tool to use here is the set of principles governing how we count atoms in a chemical formula in the preceding section. To review, the subscript following an element indicates how many of that element are part of the formula; a subscript of 1 is implied if there is no subscript. We also must recall that a quantity within parentheses is repeated a number of times equal to the subscript that follows, and a raised dot in a formula indicates the substance is a hydrate in which the number preceding H_2O specifies how many water molecules are present.

SOLUTION: (a) Here we must recognize that all the atoms within the parentheses occur three times.

Subscript 3 indicates three SO_4 units.

$$Al_2(SO_4)_3$$

Each SO_4 contains one S and four O atoms, so three of them contain three S and twelve O atoms. The subscript for Al tells us there are two Al atoms. Therefore, the formula $Al_2(SO_4)_3$ shows

2 Al 3 S 12 O

(b) This is a formula for a hydrate, as indicated by the raised dot. It contains six water molecules, each with two H and one O, for every $CoCl_2$.

The 6 indicates there are six molecules of H_2O.

$$CoCl_2{\cdot}6H_2O$$

Dot indicates the compound is a hydrate.

Therefore, the formula $CoCl_2{\cdot}6H_2O$ represents

1 Co 2 Cl 12 H 6 O

ARE THE ANSWERS REASONABLE? The only way to check the answer here is to perform a recount.

Practice Exercise 9: How many atoms of each element are expressed by the formulas below? (Hint: Pay special attention to counting elements within parentheses.)

(a) $NiCl_2$ (b) $FeSO_4$ (c) $Ca_3(PO_4)_2$ (d) $Co(NO_3)_2{\cdot}6H_2O$

Practice Exercise 10: How many atoms of each element are present in each of the formulas that follow? Consult the table inside the front cover to write the full name of each element as well as its symbol.

(a) NH_4NO_3 (b) $FeNH_4(SO_4)_2$ (c) $Mo(NO_3)_2{\cdot}5H_2O$ (d) $C_6H_4ClNO_2$

Chemical equations describe what happens in chemical reactions

A **chemical equation** *describes what happens when a chemical reaction occurs.* It uses chemical formulas to provide a before-and-after picture of the chemical substances involved. Consider, for example, the reaction between hydrogen and oxygen to give water. The chemical equation that describes this reaction is

$$2H_2 + O_2 \longrightarrow 2H_2O$$

The two substances that appear to the left of the arrow are the **reactants;** they are the substances present before the reaction begins. To the right of the arrow we find the formula for the **product** of the reaction, water. In this example, only one substance is formed in the reaction, so there is only one product. As we will see, however, in most chemical reactions there is more than one product. The products are the substances that are formed and that exist after the reaction is over. The arrow means "reacts to yield." Thus, this equation tells us that *hydrogen and oxygen react to yield water.*

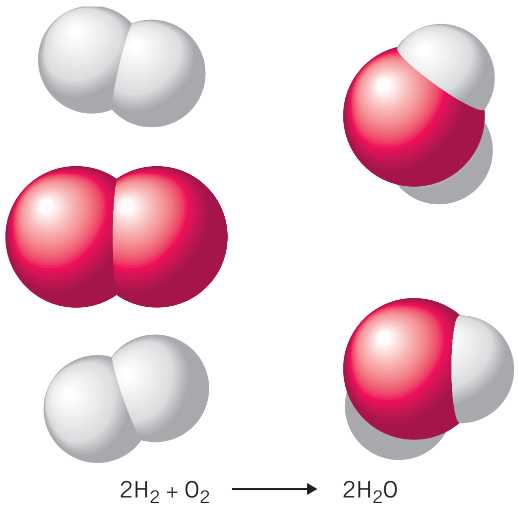


FIG. 2.18 The reaction between molecules of hydrogen and oxygen. The reaction between two molecules of hydrogen and one molecule of oxygen gives two molecules of water.

Coefficients are written in front of formulas to satisfy the law of conservation of mass

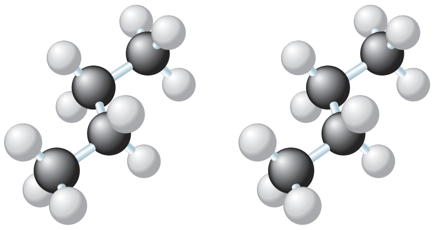
In the equation for the reaction of hydrogen and oxygen, you'll notice that the number 2 precedes the formulas of hydrogen and water. Numbers in front of the formulas are called **coefficients,** and they indicate the number of molecules of each kind among the reactants and products. Thus, $2H_2$ means two molecules of H_2, and $2H_2O$ means two molecules of H_2O. When no number is written, the coefficient is assumed to be 1 (so the coefficient of O_2 equals 1).

Coefficients are needed to have the equation conform to the law of conservation of mass, as illustrated in Figure 2.18. Because atoms cannot be created or destroyed in a chemical reaction, we must have the same number of atoms of each kind present before and after the reaction (that is, on both sides of the arrow). When this condition is met, we say the equation is **balanced.** Another example is the equation for the combustion of butane, C_4H_{10}, the fluid in disposable cigarette lighters (Figure 2.19).

Two molecules of C_4H_{10} | Thirteen molecules of O_2 | Eight molecules of CO_2 | Ten molecules of H_2O

$$2C_4H_{10} + 13O_2 \longrightarrow 8CO_2 + 10H_2O$$

FIG. 2.19 The combustion of butane, C_4H_{10}. The products are carbon dioxide and water vapor. *(Robert Capece.)*



Two molecules of butane contain 8 atoms of C and 20 atoms of H.

FIG. 2.20 Understanding coefficients in an equation. The expression $2C_4H_{10}$ describes two molecules of butane, each of which contains 4 carbon and 10 hydrogen atoms. This gives a total of 8 carbon and 20 hydrogen atoms.

The 2 before the C_4H_{10} tells us that two molecules of butane react. This involves a total of 8 carbon atoms and 20 hydrogen atoms, as we see in Figure 2.20. Notice we have multiplied the numbers of atoms of C and H in one molecule of C_4H_{10} by the coefficient 2. On the right we find 8 molecules of CO_2, which contain a total of 8 carbon atoms. Similarly, 10 water molecules contain 20 hydrogen atoms. Finally, we can count 26 oxygen atoms on both sides of the equation. You will learn to balance equations such as this in Chapter 3.

The states of the reactants and products can be specified in a chemical equation

In a chemical equation we sometimes find it useful to specify the physical states of the reactants and products, that is, whether they are solids, liquids, or gases. This is done by writing *s* for solid, *l* for liquid, or *g* for gas in parentheses after the chemical formulas. For example, the equation for the combustion of the carbon in a charcoal briquette can be written as

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

At times we will also find it useful to indicate that a particular substance is dissolved in water. We do this by writing *aq*, meaning "*aqueous* solution," in parentheses after the formula. For instance, the reaction between stomach acid (an aqueous solution of HCl) and the active ingredient in Tums, $CaCO_3$, is

$$2HCl(aq) + CaCO_3(s) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Remember:

s = solid
l = liquid
g = gas
aq = aqueous

EXAMPLE 2.5
Determining if an Equation Is Balanced

Determine whether or not the following chemical equations are balanced. Support your conclusions by writing how many atoms of each element are on each side of the arrow.

(a) $Fe(OH)_3 + 2HNO_3 \longrightarrow Fe(NO_3)_3 + 2H_2O$

(b) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$

(c) $C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O$

ANALYSIS: The statement of the problem asks if the equations are balanced. You can prove an equation is balanced if each element has the same number of atoms on each side of the arrow. Again, in this example, we use the tool that tells us how to use subscripts to count atoms in each formula. Also, a given atom, such as oxygen, in these equations may appear in both reactants or both products and we need to be sure to account for all of them.

SOLUTION: (a) Reactants: 1 Fe, 9 O, 5 H, 2 N. Products: 1 Fe, 11 O, 4 H, 3 N.
Only Fe has the same number of atoms on each side of the arrow. This is *not* balanced.
(b) Reactants: 1 Ba, 2 Cl, 2 H, 1 S, 4 O. Products: 1 Ba, 2 Cl, 2 H, 1 S, 4 O.
This equation *is* balanced.
(c) Reactants: 6 C, 12 H, 18 O. Products: 6 C, 18 O, 12 H.
This equation *is* balanced.

ARE THE ANSWERS REASONABLE? The appropriate way to check this is to recount the atoms. Try counting the atoms in the reverse direction this time.

Practice Exercise 11: How many atoms of each element appear on each side of the arrow in the following equation? (Hint: Recall that coefficients multiply the elements in the entire formula.)

$$Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O$$

Practice Exercise 12: Rewrite the equation in Practice Exercise 11 to show that $Mg(OH)_2$ is a solid, HCl and $MgCl_2$ are dissolved in water, and H_2O is a liquid.

Practice Exercise 13: Count each of the atoms in the following equation to determine if it is balanced.

$$2(NH_4)_3PO_4 + 3Ba(C_2H_3O_2)_2 \longrightarrow Ba_3(PO_4)_2 + 6NH_4C_2H_3O_2$$

2.6 Molecular Compounds Contain Neutral Particles Called Molecules

The concept of molecules dates to the time of Dalton's atomic theory, where a part of his theory was that atoms of elements combine in fixed numerical ratios to form "molecules" of a compound. By our modern definition *a* **molecule** *is an electrically neutral particle consisting to two or more atoms.* Accordingly, the term molecule applies to many elements such as H_2 and O_2 as well as to **molecular compounds**.

Experimental evidence exists for molecules

One phenomenon that points to the existence of molecules is called **Brownian motion** [named after Robert Brown (1773–1858), the Scottish botanist who first observed it]. When very small particles such as tiny grains of pollen are suspended in a liquid and observed under a microscope, the tiny particles are seen to be constantly jumping and jiggling about. It appears as though they are continually being knocked back and forth by collisions with something. An explanation is that this "something" is *molecules* of the liquid. The microscopic particles are constantly bombarded by molecules of the liquid, but because the suspended particles are so small, the collisions are not occurring equally on all sides. The unequal numbers of collisions cause the lightweight particles to jerk about.

There is additional evidence for molecules, and today scientists accept the existence of molecules as fact. Looking more closely, within molecules atoms are held to each other by attractions called **chemical bonds,** which are electrical in nature. In molecular compounds chemical bonds arise from the sharing of electrons between one atom and another. We will discuss such bonds at considerable length in Chapters 8 and 9. What is important to know about molecules now is that *the group of atoms that make up a molecule move about together and behave as a single particle*, just as the various parts that make up a car move about as one unit. The chemical formulas that we write to describe the compositions of molecules are called **molecular formulas,** which specify the actual numbers of atoms of each kind that make up a single molecule.

Molecular compounds form when nonmetals combine

As a general rule, *molecular compounds are formed when nonmetallic elements combine.* For example, you learned that H_2 and O_2 combine to form molecules of water. Similarly, carbon and oxygen combine to form either carbon monoxide, CO, or carbon dioxide, CO_2. (Both are gases that are formed in various amounts as products in the combustion of fuels such as gasoline and charcoal.) Although molecular compounds can be formed by the direct combination of elements, often they are the products of reactions between compounds. You will encounter many such reactions in your study of chemistry.

Carbon monoxide is a poisonous gas found in the exhaust of automobiles.

Although there are relatively few nonmetals, the number of molecular substances formed by them is huge. This is because of the variety of ways in which they combine as well as the varying degrees of complexity of their molecules. Variety and complexity reach a maximum with compounds in which carbon is combined with a handful of other elements such as hydrogen, oxygen, and nitrogen. There are so many of these compounds, in fact, that their study encompasses the chemical specialties called organic chemistry and biochemistry.

Molecules vary in size from small to very large. Some contain as few as two atoms (diatomic molecules). Most molecules are more complex, however, and contain more atoms.

Molecules of water (H_2O), for example, have three atoms and those of ordinary table sugar ($C_{12}H_{22}O_{11}$) have 45. There also are molecules that are very large, such as those that occur in plastics and in living organisms, some of which contain millions of atoms.

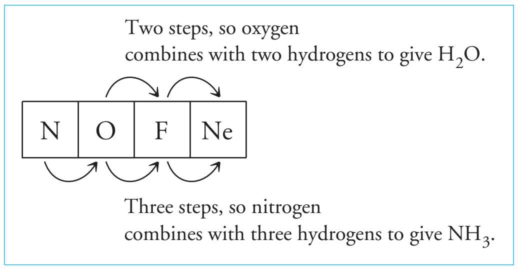
At this early stage we can only begin to look for signs of order among the vast number of nonmetal–nonmetal compounds. To give you a taste of the subject, we will look briefly at some simple compounds that the nonmetals form with hydrogen, as well as some simple compounds of carbon.

Hydrogen forms compounds with many nonmetals

Compounds that elements form with hydrogen are often called **hydrides,** and the formulas of the simple hydrides of the nonmetals are given in Table 2.4.[1] These compounds provide an opportunity to observe how we can use the periodic table as an aid in remembering factual information, in this case, the formulas of the hydrides. Notice that the number of hydrogen atoms combined with the nonmetal atom equals *the number of spaces to the right that we have to move in the periodic table to get to a noble gas.* (You will learn *why* this is so in Chapter 8, but for now we can just use the periodic table to help us remember the formulas.)

TOOLS

Predicting hydride formulas



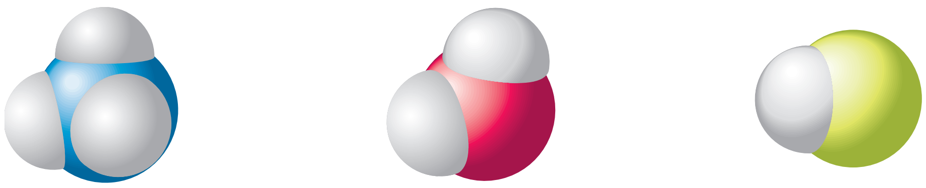
Many of the nonmetals form more complex compounds with hydrogen, but we will not discuss them here.

Also note in Table 2.4 that the formulas of the simple hydrides are similar for nonmetals within a given group of the periodic table. If you know the formula for the hydride of the top member of the group, then you know the formulas of all of them in that group.

We live in a three-dimensional world, and this is reflected in the three-dimensional shapes of molecules. The shapes of the simple nonmetal hydrides of nitrogen, oxygen, and fluorine are illustrated as space-filling models in Figure 2.21. Our understanding of the geometric shapes of molecules is described in Chapter 9.

TABLE 2.4 Simple Hydrogen Compounds of the Nonmetallic Elements

	Group			
Period	IVA	VA	VIA	VIIA
2	CH_4	NH_3	H_2O	HF
3	SiH_4	PH_3	H_2S	HCl
4	GeH_4	AsH_3	H_2Se	HBr
5		SbH_3	H_2Te	HI



Ammonia, NH_3 Water, H_2O Hydrogen fluoride, HF

FIG. 2.21 Nonmetal hydrides of nitrogen, oxygen, and fluorine.

[1] Table 2.4 shows how the formulas are normally written. The order in which the hydrogens appear in the formula is not of concern to us now. Instead, we are interested in the *number* of hydrogens that combine with a given nonmetal.

Compounds of carbon form the basis for organic chemistry

Among all the elements, carbon is unique in the variety of compounds it forms with elements such as hydrogen, oxygen, and nitrogen. As a consequence, the number and complexity of such **organic compounds** is enormous, and their study constitutes the major specialty called **organic chemistry**. The term *organic* here comes from an early belief that these compounds could only be made by living organisms. We now know this isn't true, but the name organic chemistry persists nonetheless.

Our goal at this time is to acquaint you with some of the important kinds of organic compounds we encounter regularly, so our discussion here is brief.

Organic compounds are around us everywhere and we will frequently use such substances as examples in our discussions. Therefore, it will be helpful if you can begin to learn some of them now.

The study of organic chemistry begins with **hydrocarbons** (compounds of carbon and hydrogen). The simplest hydrocarbon is methane, CH_4, which is a member of a series of hydrocarbons with the general formula C_nH_{2n+2}, where n is an integer (i.e., a whole number). The first six members of this series, called the **alkane** series, are given in Table 2.5 along with their boiling points. Notice that as the molecules become larger, their boiling points increase.

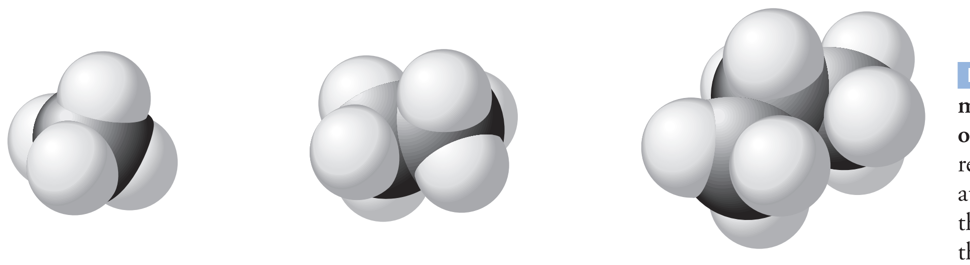
TABLE 2.5 Hydrocarbons Belonging to the Alkane Series

Compound	Name	Boiling Point (°C)
CH_4	Methane[a]	−161.5
C_2H_6	Ethane[a]	−88.6
C_3H_8	Propane[a]	−42.1
C_4H_{10}	Butane[a]	−0.5
C_5H_{12}	Pentane	36.1
C_6H_{14}	Hexane	68.7

[a] Gases at room temperature (25 °C) and atmospheric pressure.

Molecular formulas can be written in different ways depending on what information is needed. A condensed formula such as C_2H_6 for ethane or C_3H_8 for propane simply indicates the number of each type of atom in the molecule. Structural formulas indicate how the carbon atoms are connected. Ethane is written as CH_3CH_3 and propane is $CH_3CH_2CH_3$ in the structural format. Line structures and a ball-and-stick representation are illustrated in Figure 1.3 for methane, while Figure 2.22. illustrates ethane and propane in space-filling models.

The alkanes are common substances. They are the principal constituents of petroleum from which most of our useful fuels are produced. Methane itself is the major component of natural gas that is often used for home heating and cooking. Gas-fired barbecues and some homes use propane as a fuel, and butane is the fuel in inexpensive cigarette lighters.[2] Hydrocarbons with higher boiling points are found in gasoline, kerosene, paint thinners, diesel fuel, and even candle wax.



Methane, CH_4 Ethane, C_2H_6 Propane, C_3H_8

FIG. 2.22 The first three members of the alkane series of hydrocarbons. White atoms represent hydrogen and black atoms represent carbon in these space-filling models that demonstrate the shapes of the molecules.

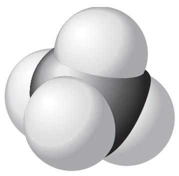
[2] Propane and butane are gases when they're at the pressure of the air around us, but they become liquids when compressed. When you purchase these substances, they are liquids with pressurized gas above them. The gas can be drawn off and used by opening a valve to the container.

Alkanes are not the only class of hydrocarbons. For example, there are three two-carbon hydrocarbons. In addition to ethane, C_2H_6, there are ethylene, C_2H_4 (an alkene from which polyethylene is made), and acetylene, C_2H_2 (an alkyne, which is the fuel used in *acetylene* welding torches).

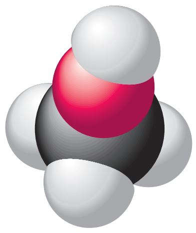
The chemically correct names for ethylene and acetylene are ethene and ethyne, respectively.

The hydrocarbons serve as the foundation for organic chemistry. Derived from them are various other classes of organic compounds. An example is the class of compounds called **alcohols,** in which the atoms OH replace a hydrogen in the hydrocarbon. Thus, *methanol*, CH_3OH (also called *methyl alcohol*), is related to methane, CH_4, by removing one H and replacing it with OH (Figure 2.23). Methanol is used as a fuel and as a raw material for making other organic chemicals. Another familiar alcohol is *ethanol* (also called *ethyl alcohol*), C_2H_5OH. Ethanol, known as grain alcohol because it is obtained from the fermentation of grains, is in alcoholic beverages. It is also mixed with gasoline to reduce petroleum consumption. A 10% ethanol/90% gasoline mixture is known as gasohol; an 85% mixture of ethanol and gasoline is called E85.

Methanol is also known as wood alcohol because it was originally made by distilling wood. It is quite poisonous. Ethanol in high doses is also a poison.



Methane



Methanol

FIG. 2.23 Relationship between an alkane and an alcohol. The alcohol methanol is derived from methane by replacing one H by OH. (Color code: carbon is black, hydrogen is white, and oxygen is red.)

Alcohols constitute just one class of compounds derived from hydrocarbons. We will discuss some others after you've learned more about how atoms bond to each other and about the structures of molecules.

Practice Exercise 14: Gasoline used in modern cars is a complex mixture of hundreds of different organic compounds. Less than 1% of gasoline is actually octane. Write the formula for octane using the condensed and structural format. (Hint: The prefix "octa-" stands for the number 8.)

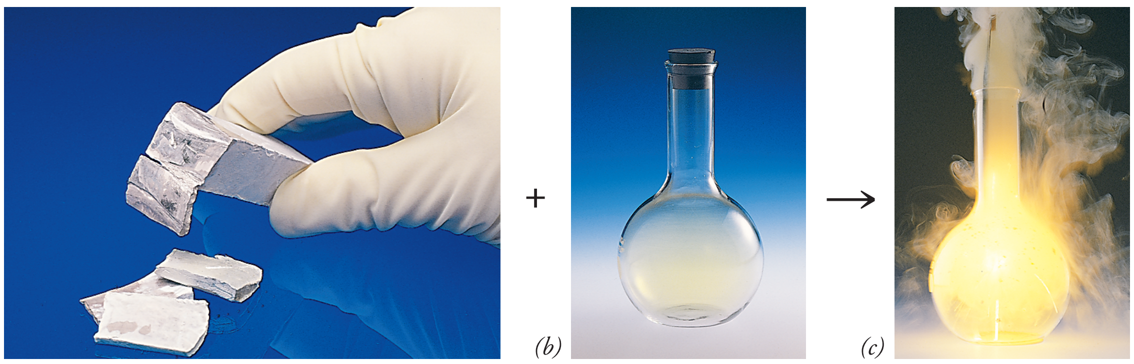
Practice Exercise 15: What is the formula of the alkane hydrocarbon having 10 carbon atoms, decane? Write the formula in the condensed form and in the structural form. Download the space-filling molecule from the Internet.

Practice Exercise 16: On the basis of the discussions in this section, what are the formulas of (a) propanol and (b) butanol? Write the formula in the condensed form and in the structural form. Download the space-filling molecules from the Internet.

2.7 Ionic compounds are composed of charged particles called ions

Formation of an ionic compound from its elements involves electron transfer

Under appropriate conditions, atoms are able to transfer electrons between one another when they react. This is what happens, for example, when the metal sodium combines with the nonmetal chlorine. As shown in Figure 2.24, sodium is a typical shiny metal and chlorine is



(a) + *(b)* → *(c)*

FIG. 2.24 Sodium reacts with chlorine to give the ionic compound sodium chloride. (*a*) Freshly cut sodium has a shiny metallic surface. The metal reacts with oxygen and moisture, so it cannot be touched with bare fingers. (*b*) Chlorine is a pale green gas. (*c*) When a small piece of sodium is melted in a metal spoon and thrust into the flask of chlorine, it burns brightly as the two elements react to form sodium chloride. The smoke coming from the flask is composed of fine crystals of salt. *(Michael Watson; Richard Megna/Fundamental Photographs; Richard Megna/Fundamental Photographs.)*

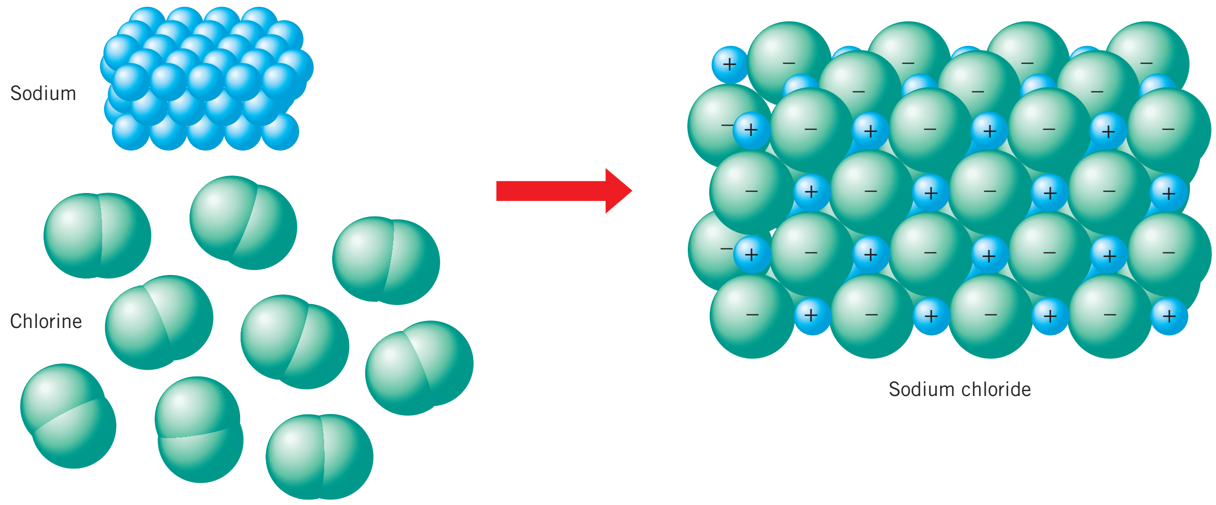


FIG. 2.25 The reaction of sodium with chlorine viewed at the atomic level. Electrically neutral atoms and molecules react to yield positive and negative ions which are held to each other by electrostatic attractions (attractions between opposite electrical charges).

a pale green gas. When a piece of heated sodium is thrust into the chlorine, a vigorous reaction takes place yielding a white powder, salt (NaCl). The equation for the reaction is

$$2\mathrm{Na}(s) + \mathrm{Cl}_2(g) \longrightarrow 2\mathrm{NaCl}(s)$$

The changes that take place at the atomic level are illustrated in Figure 2.25.

The formation of the **ions** in salt results from the transfer of electrons between the reacting atoms. Specifically, each sodium atom gives up one electron to a chlorine atom. We can diagram the changes in equation form by using the symbol e^- to stand for an electron.

$$\overset{e^-}{\mathrm{Na} + \mathrm{Cl}} \longrightarrow \mathrm{Na}^+ + \mathrm{Cl}^-$$

The electrically charged particles formed in this reaction are a sodium ion (Na^+) and a chloride ion (Cl^-). The sodium ion has a positive 1+ charge, indicated by the superscript plus sign, because the loss of an electron leaves it with one more proton in its nucleus than there are electrons outside. Similarly, by gaining one electron the chlorine atom has added one more negative charge, so the chloride ion has a single negative charge indicated by the minus sign. Solid sodium chloride is composed of these charged sodium and chloride ions and is said to be an **ionic compound.**

As a general rule, *ionic compounds are formed when metals react with nonmetals*. In the electron transfer, however, not all atoms gain or lose just one electron; some gain or lose more. For example, when calcium atoms react, they lose two electrons to form Ca^{2+} ions and when oxygen atoms form ions they each gain two electrons to give O^{2-} ions. (We will have to wait until a later chapter to study the reasons why certain atoms gain or lose one electron each, whereas other atoms gain or lose two or more electrons.)

Here we are concentrating on what happens to the individual atoms, so we have not shown chlorine as diatomic Cl_2 molecules.

A neutral sodium atom has 11 protons and 11 electrons; a sodium ion has 11 protons and 10 electrons, so it carries a unit positive charge. A neutral chlorine atom has 17 protons and 17 electrons; a chloride ion has 17 protons and 18 electrons, so it carries a unit negative charge.

Practice Exercise 17: For each of the following atoms or ions, give the number of protons and the number of electrons in one particle. (a) an Fe atom, (b) an Fe^{3+} ion, (c) an N^{3-} ion, (d) an N atom. (Hint: Recall that electrons have a negative charge and ions that have a negative charge are atoms must have gained electrons.)

Practice Exercise 18: For each of the following atoms or ions, give the number of protons and the number of electrons in one particle. (a) an O atom, (b) an O^{2-} ion, (c) an Al^{3+} ion, (d) an Al atom.

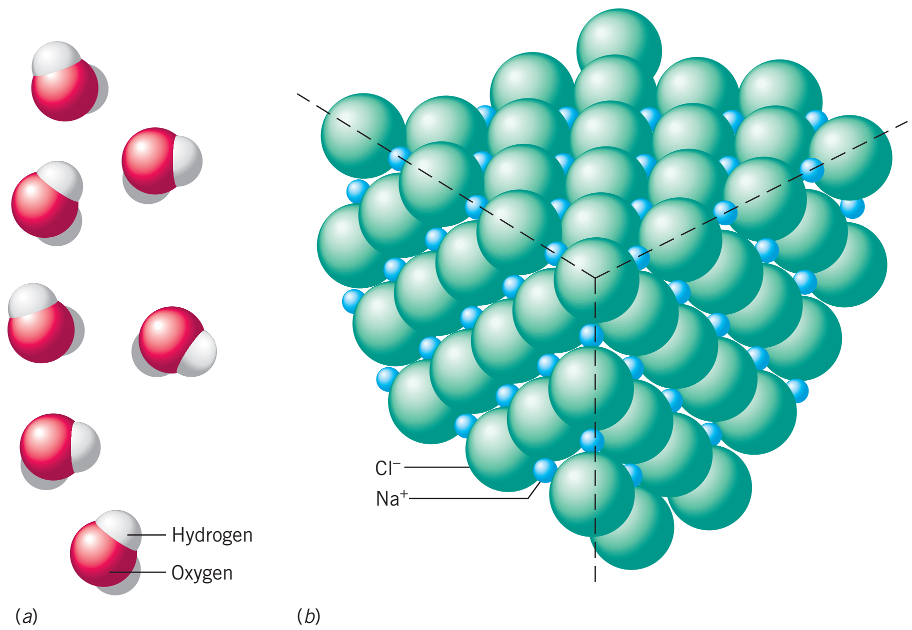


FIG. 2.26 **Molecular and ionic substances.** (*a*) In water there are discrete molecules that each consist of one atom of oxygen and two atoms of hydrogen. Each particle has the formula H_2O. (*b*) In sodium chloride, ions are packed in the most efficient way. Each Na^+ is surrounded by six Cl^-, and each Cl^- is surrounded by six Na^+. Because individual molecules do not exist, we simply specify the ratio of ions as NaCl.

Notice that the charges on the ions are omitted when writing formulas for compounds. This is because compounds are electrically neutral overall.

Figure 2.26 compares the structures of water and sodium chloride and demonstrates an important difference between molecular and ionic compounds. In water it is safe to say that two hydrogen atoms "belong" to each oxygen atom in a particle having the formula H_2O. However, in NaCl it is impossible to say that a particular Na^+ ion belongs to a particular Cl^- ion. The ions in a crystal of NaCl are simply packed in the most efficient way, so that positive ions and negative ions can be as close to each other as possible. In this way, the attractions between oppositely charged ions, which are responsible for holding the compound together, can be as strong as possible.

Because molecules don't exist in ionic compounds, the subscripts in their formulas are always chosen to specify the smallest whole-number ratio of the ions. This is why the formula of sodium chloride is given as NaCl rather than Na_2Cl_2 or Na_3Cl_3. Although the smallest unit of an ionic compound can't be called a molecule, the idea of "smallest unit" is still quite often useful. Therefore, we take the smallest unit of an ionic compound to be whatever is represented in its formula and call this unit a **formula unit.** Thus, one formula unit of NaCl consists of one Na^+ and one Cl^-, whereas one formula unit of the ionic compound $CaCl_2$ consists of one Ca^{2+} and two Cl^- ions. (In a broader sense, we can use the term *formula unit* to refer to whatever is represented by a formula. Sometimes the formula specifies a set of ions, as in NaCl; sometimes it is a molecule, as in O_2 or H_2O; sometimes it can be just an ion, as in Cl^- or Ca^{2+}; and sometimes it might be just an atom, as in Na.)

Experimental evidence exists for ions in compounds

We know that metals conduct electricity because electrons can move from one atom to the next in a wire when connected to a battery. Solid ionic compounds are poor conductors of electricity as are molecular substances such as water. However, if an ionic compound is dissolved in water or is heated to a high temperature, so that it melts, the resulting liquids are able to conduct electricity easily. These observations suggest that ionic compounds are composed of charged ions rather than neutral molecules and these ions when made mobile by dissolving or melting can conduct electricity. Figure 2.27 illustrates how the electrical conductivity can be tested.

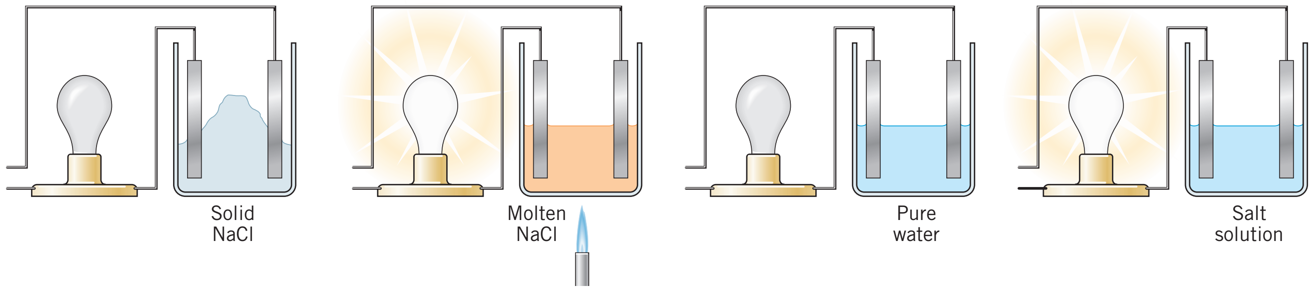


FIG. 2.27 **An apparatus to test for electrical conductivity.** The electrodes are dipped into the substance to be tested. If the lightbulb glows when electricity is applied, the sample is an electrical conductor. Here we see that solid sodium chloride does not conduct electricity, but when the solid is melted it does conduct. Liquid water, a molecular compound, is not a conductor of electricity because it does not contain electrically charged particles. An aqueous salt solution contains ions of the salt and will conduct electricity.

2.8 The Formulas of Many Ionic Compounds Can Be Predicted

In the preceding section we noted that metals combine with nonmetals to form ionic compounds. In such reactions, metal atoms lose one or more electrons to become positively charged ions and nonmetal atoms gain one or more electrons to become negatively charged ions. In referring to these particles, we will frequently call a positively charged ion a **cation** (pronounced *CAT-i-on*) and a negatively charged ion an **anion** (pronounced *AN-i-on*).[3] Thus, solid NaCl is composed of sodium cations and chloride anions.

Ions formed by representative metals and nonmetals can be remembered using the periodic table

In Section 2.6 you saw that the periodic table can be helpful in remembering the formulas of the nonmetal hydrides. It can also help us remember the kinds of ions formed by many of the representative elements (elements in the A groups of the periodic table). For example, except for hydrogen, the neutral atoms of the Group IA elements always lose one electron each when they react, thereby becoming ions with a charge of 1+. Similarly, atoms of the Group IIA elements always lose two electrons when they react; so these elements always form ions with a charge of 2+. In Group IIIA, the only important positive ion we need consider now is that of aluminum, Al^{3+}; an aluminum atom loses three electrons when it reacts to form the ion.

Positive ions, called **cations**, are usually metals that have lost one or more electrons.

TOOLS
Predicting cation charge

All these ions are listed in Table 2.6. *Notice that the number of positive charges on each of the cations is the same as the group number when we use the North American numbering of groups in the periodic table.* Thus, sodium is in Group IA and forms an ion with a 1+

TABLE 2.6 **Some Ions Formed from the Representative Elements**

Group Number						
IA	IIA	IIIA	IVA	VA	VIA	VIIA
Li^+	Be^{2+}		C^{4-}	N^{3-}	O^{2-}	F^-
Na^+	Mg^{2+}	Al^{3+}	Si^{4-}	P^{3-}	S^{2-}	Cl^-
K^+	Ca^{2+}				Se^{2-}	Br^-
Rb^+	Sr^{2+}				Te^{2-}	I^-
Cs^+	Ba^{2+}					

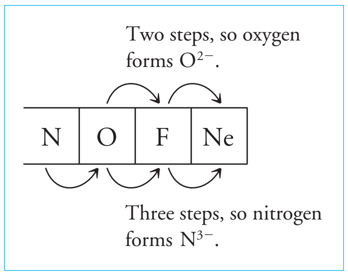
[3] The names *cation* and *anion* come from the way the ions behave when electrically charged metal plates called electrodes are dipped into a solution that contains them. We will discuss this in detail in Chapter 19.

charge, barium (Ba) is in Group IIA and forms an ion with a 2+ charge, and aluminum is in Group IIIA and forms an ion with a 3+ charge. Although this generalization doesn't work for all the metallic elements (it doesn't work for the transition elements, for instance), it does help us remember what happens to the metallic elements of Groups IA and IIA and aluminum when they react.

Negative ions, called **anions**, are monatomic nonmetals that have gained one or more electrons.

Among the nonmetals on the right side of the periodic table we also find some useful generalizations. For example, when they combine with metals, the halogens (Group VIIA) form ions with one negative charge (written as 1−) and the nonmetals in Group VIA form ions with two negative charges (written as 2−). Notice that *the number of negative charges on the anion is equal to the number of spaces to the right that we have to move in the periodic table to get to a noble gas.*

TOOLS
Predicting anion charge



You've probably also noticed that the number of negative charges on an anion is the same as the number of hydrogens in the simple nonmetal hydride of the element, as shown in Table 2.4.

Writing formulas for ionic compounds follows certain rules

All chemical compounds are electrically neutral, so the ions in an ionic compound always occur in a ratio such that the total positive charge is equal to the total negative charge. This is why the formula for sodium chloride is NaCl; the 1-to-1 ratio of Na^+ to Cl^- gives electrical neutrality. In addition, as we've already mentioned, discrete molecules do not exist in ionic compounds, so we always use the smallest set of subscripts that specify the correct ratio of the ions. The following, therefore, are the rules we use in writing the formulas of ionic compounds.

TOOLS
Formulas for ionic compounds

A substance is electrically neutral, with a net charge of zero, if the total positive charge equals the total negative charge.

Rules for Writing Formulas of Ionic Compounds

1. The positive ion is given first in the formula. (This isn't required by nature, but it is a custom we always follow.)
2. The subscripts in the formula must produce an electrically neutral formula unit. (Nature *does* require electrical neutrality.)
3. The subscripts should be the smallest set of whole numbers possible. For instance, if all subscripts are even, divide them by 2. (You may have to repeat this simplification step several times.)
4. The charges on the ions are not included in the finished formula for the substance. When a subscript is 1 it is left off; no subscript implies a subscript of 1.

EXAMPLE 2.6
Writing Formulas for Ionic Compounds

Write the formulas for the ionic compounds formed from (a) Ba and S, (b) Al and Cl, and (c) Al and O.

ANALYSIS: To correctly write the formula, we have to apply our tool that summarizes the rules for ionic compounds listed above. First, we need to figure out the charges of the ions and since we're working with representative elements, we can use the periodic table to do this. Then we need to assemble the ions so that the formula unit is electrically neutral.

SOLUTION: (a) The element Ba is in Group IIA, so the charge on its ion is 2+. Sulfur is in Group VIA, so its ion has a charge of 2−. Therefore, the ions are Ba^{2+} and S^{2-}. Since the charges are equal but opposite, a 1-to-1 ratio will give a neutral formula unit. Therefore, the formula is BaS. Notice that we have not included the charges on the ions in the finished formula.

(b) By using the periodic table, the ions of these elements are Al^{3+} and Cl^-. We can obtain a neutral formula unit by combining one Al^{3+} with three Cl^-. (The charge on Cl^- is 1−; the 1 is understood.)

$$1(3+) + 3(1-) = 0$$

The formula is $AlCl_3$.

(c) For these elements, the ions are Al^{3+} and O^{2-}. In the formula we seek there must be the same number of positive charges as negative charges. This number must be a whole-number multiple of both 3 and 2. The smallest number that satisfies this condition is 6, so there must be two Al^{3+} and three O^{2-} in the formula.

$$\begin{array}{ll} 2Al^{3+} & 2(3+) = 6+ \\ 3O^{2-} & \underline{3(2-) = 6-} \\ & \quad \text{sum} = 0 \end{array}$$

The formula is Al_2O_3.

A "trick" you may have seen before is to use the *number* of positive charges for the subscript of the anion and the *number* of negative charges as the subscript for the cation as shown in the diagram.

$$\mathrm{Al}^{3+} \qquad \mathrm{O}^{2-}$$

When using this method, always be sure to check that the subscripts cannot be reduced to smaller numbers.

ARE THE ANSWERS REASONABLE? In writing a formula, there are two things to check. First, be sure you've correctly written the formulas of the ions. (This is where students make a lot of mistakes.) Then check that you've combined them in a ratio that gives electrical neutrality.

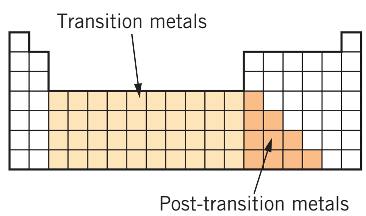
Practice Exercise 19: Write formulas for ionic compounds formed from (a) Na and F, (b) Na and O, (c) Mg and F, and (d) Al and C. (Hint: One element must form a cation and the other will form an anion based on its position in the periodic table.)

Practice Exercise 20: Write the formulas for the compounds made from (a) Ca and N, (b) Al and Br, (c) Na and P, and (d) Cs and Cl.

Many of our most important chemicals are ionic compounds. We have mentioned NaCl, common table salt, and $CaCl_2$, which is a substance often used to melt ice on walkways in the winter. Other examples are sodium fluoride, NaF, used by dentists to give fluoride treatments to teeth, and calcium oxide, CaO, an important ingredient in cement.

Transition and post-transition metals form more than one cation

The transition elements are located in the center of the periodic table, from Group IIIB on the left to Group IIB on the right (Groups 3 to 12 using the IUPAC system). All of them lie to the left of the metalloids, and they all are metals. Included here are some of our most familiar metals, including iron, chromium, copper, silver, and gold.



Distribution of transition and post-transition metals in the periodic table.

TABLE 2.7 Ions of Some Transition Metals and Post-transition Metals

Transition Metals	
Chromium	Cr^{2+}, Cr^{3+}
Manganese	Mn^{2+}, Mn^{3+}
Iron	Fe^{2+}, Fe^{3+}
Cobalt	Co^{2+}, Co^{3+}
Nickel	Ni^{2+}
Copper	Cu^{+}, Cu^{2+}
Zinc	Zn^{2+}
Silver	Ag^{+}
Cadmium	Cd^{2+}
Gold	Au^{+}, Au^{3+}
Mercury	Hg_2^{2+}, Hg^{2+}
Post-transition Metals	
Tin	Sn^{2+}, Sn^{4+}
Lead	Pb^{2+}, Pb^{4+}
Bismuth	Bi^{3+}

Most of the transition metals are much less reactive than the metals of Groups IA and IIA, but when they react they also transfer electrons to nonmetal atoms to form ionic compounds. However, the charges on the ions of the transition metals do not follow as straightforward a pattern as do those of the alkali and alkaline earth metals. One of the characteristic features of the transition metals is the ability of many of them to form more than one positive ion. Iron, for example, can form two different ions, Fe^{2+} and Fe^{3+}. This means that iron can form more than one compound with a given nonmetal. For example, with the chloride ion, Cl^-, iron forms two compounds, with the formulas $FeCl_2$ and $FeCl_3$. With oxygen, we find the compounds FeO and Fe_2O_3. As usual, we see that the formulas contain the ions in a ratio that gives electrical neutrality. Some of the most common ions of the transition metals are given in Table 2.7. Notice that one of the ions of mercury is diatomic Hg_2^{2+}. It consists of two Hg^+ ions joined by the same kind of bond found in molecular substances. The simple Hg^+ ion does not exist.

The prefix *post* means "after."

The **post-transition metals** are those metals that occur in the periodic table immediately following a row of transition metals. The two most common and important ones are tin (Sn) and lead (Pb). Except for bismuth, post-transition metals have the ability to form two different ions, and therefore two different compounds with a given nonmetal. For example, tin forms two oxides, SnO and SnO_2. Lead also forms two oxides that have similar formulas (PbO and PbO_2). The ions that these metals form are also included in Table 2.7.

Practice Exercise 21: Write formulas for the chlorides and oxides formed by (a) chromium and (b) copper. (Hint: There are more than one chloride and oxide for each of these transition metals.)

Practice Exercise 22: Write the formulas for the sulfides and nitrides of (a) gold and (b) tin.

Ions may be composed of more than one element

A substance is **diatomic** if it is composed of molecules that contain only two atoms. It is a **binary compound** if it contains two different elements, regardless of the number of each. Thus, BrCl is a binary compound and is also diatomic; CH_4 is a binary compound but is not diatomic.

The metal compounds that we have discussed so far have been **binary compounds**—compounds formed from *two* different elements. There are many other ionic compounds that contain more than two elements. These substances usually contain **polyatomic ions,** which are ions that are themselves composed of two or more atoms linked by the same kinds of bonds that hold molecules together. Polyatomic ions differ from molecules, however, in that they contain either too many or too few electrons to make them electrically neutral. Table 2.8 lists some important polyatomic ions. It is very important that you learn the formulas, charges, and names of all of these ions.

The formulas of compounds formed from polyatomic ions are determined in the same way as are those of binary ionic compounds; the ratio of the ions must be such that the formula unit is electrically neutral, and the smallest set of whole-number subscripts is used. One difference in writing formulas with polyatomic ions is that parentheses are needed around the polyatomic ion if a subscript is required.

TABLE 2.8 Formulas and Names of Some Polyatomic Ions

Ion	Name (Alternate Name in Parentheses)
NH_4^+	ammonium ion
H_3O^+	hydronium ion[a]
OH^-	hydroxide ion
CN^-	cyanide ion
NO_2^-	nitrite ion
NO_3^-	nitrate ion
ClO^- or OCl^-	hypochlorite ion
ClO_2^-	chlorite ion
ClO_3^-	chlorate ion
ClO_4^-	perchlorate ion
MnO_4^-	permanganate ion
$C_2H_3O_2^-$	acetate ion
$C_2O_4^{2-}$	oxalate ion
CO_3^{2-}	carbonate ion
HCO_3^-	hydrogen carbonate ion (bicarbonate ion)[b]
SO_3^{2-}	sulfite ion
HSO_3^-	hydrogen sulfite ion (bisulfite ion)[b]
SO_4^{2-}	sulfate ion
HSO_4^-	hydrogen sulfate ion (bisulfate ion)[b]
SCN^-	thiocyanate ion
$S_2O_3^{2-}$	thiosulfate ion
CrO_4^{2-}	chromate ion
$Cr_2O_7^{2-}$	dichromate ion
PO_4^{3-}	phosphate ion
HPO_4^{2-}	monohydrogen phosphate ion
$H_2PO_4^-$	dihydrogen phosphate ion

[a]You will encounter this ion only in aqueous solutions.
[b]You will often see and hear the alternate names for these ions.

TOOLS
Polyatomic ions

In general, polyatomic ions are not formed by the direct combination of elements. They are the products of reactions between compounds.

EXAMPLE 2.7
Formulas That Contain Polyatomic Ions

One of the minerals responsible for the strength of bones is the ionic compound calcium phosphate, which is formed from Ca^{2+} and PO_4^{3-}. Write the formula for this compound.

ANALYSIS: The essential tool for solving this problem is the identity of the formula, including the charge, of the polyatomic ion. We have related much information about ions to the periodic table. Unfortunately, the polyatomic ions must be memorized. Knowledge of the polyatomic ions is required for this problem.

SOLUTION: As before, if the number of positive charges on the cation is equal to the number of negative of charges on the anion, the formula unit will contain one of each. If the number of charges are not equal then we use the number of positive charges as the subscript for the anion and the number of negative charges as the subscript for the cation. We will need three calcium ions to give a total charge of 6+ and two phosphate ions to give a charge of 6− so that the total charge is $+6 - 6 = 0$. The formula is written with parentheses to show that the PO_4^{3-} ion occurs two times in the formula unit.

$$Ca_3(PO_4)_2$$

IS THE ANSWER REASONABLE? We double-check to see that electrical neutrality is achieved for the compound. We have six positive charges from the three Ca^{2+} ions and six negative charges from the two PO_4^{3-} ions. The sum is zero and our compound is electrically neutral as required.

Practice Exercise 23: Write the formula for the ionic compound formed from (a) potassium ion and acetate ion, (b) strontium ion and nitrate ion, and (c) Fe^{3+} and acetate ion. (Hint: See if you remember these polyatomic ions before looking at the table.)

Practice Exercise 24: Write the formula for the ionic compound formed from (a) Na^+ and CO_3^{2-}, (b) NH_4^+ and SO_4^{2-}.

Polyatomic ions are found in a large number of very important compounds. Examples include $CaSO_4$ (calcium sulfate, in plaster of Paris), $NaHCO_3$ (sodium bicarbonate, also called baking soda), NaOCl (sodium hypochlorite, in liquid laundry bleach), $NaNO_2$ (sodium nitrite, a meat preservative), $MgSO_4$ (magnesium sulfate, also known as Epsom salts), and $NH_4H_2PO_4$ (ammonium dihydrogen phosphate, a fertilizer).

2.9 Molecular and Ionic Compounds Are Named Following a System

In conversation, chemists rarely use formulas to describe compounds. Instead, names are used. For example, you already know that water is the name for the compound having the formula H_2O and that sodium chloride is the name of NaCl.

At one time there was no uniform procedure for assigning names to compounds, and those who discovered compounds used whatever method they wished. Today, we know of more than 15 million different chemical compounds, and it is necessary to have a logical system for naming them. Chemists around the world now agree on a systematic method for naming substances that is overseen by the International Union of Pure and Applied Chemistry, IUPAC. By using the **IUPAC rules** we are able to write the correct formula given the name for the many compounds we will encounter. Additionally we will be able to take a formula and correctly name it.

In this section we discuss the **nomenclature** (naming) of simple molecular and ionic inorganic compounds. In general, **inorganic compounds** are substances that would *not* be considered to be derived from hydrocarbons such as methane (CH_4), ethane (C_2H_6), and other carbon–hydrogen compounds. As we noted earlier, the hydrocarbons and compounds that can be thought of as coming from them are called organic compounds. We will have more to say about naming organic compounds later.

Even if we exclude organic compounds, the number and variety of molecular substances is quite enormous. To introduce you to the naming of them, we will restrict ourselves to binary compounds.

Binary compounds composed of two nonmetals are named using Greek prefixes

Our goal is to be able to translate a chemical formula into a name that contains information that would enable someone else, just looking at the name, to reconstruct the formula. For a binary molecular compound, therefore, we must indicate which two elements are present and the number of atoms of each in a molecule of the substance.

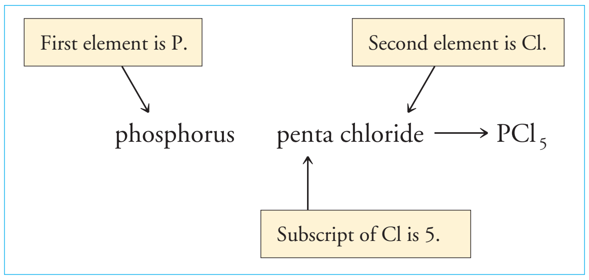
To identify the first element in a formula, we just specify its English name. Thus, for HCl the first word in the name is "hydrogen" and for PCl_5 the first word is "phosphorus." To identify the second element, we append the suffix *-ide* to the stem of the element's English name. Here are some examples:

Element	*Stem*	*Name as second element*
oxygen	ox-	oxide
sulfur	sulf-	sulfide
nitrogen	nitr-	nitride
phosphorus	phosph-	phosphide
fluorine	fluor-	fluoride
chlorine	chlor-	chloride
bromine	brom-	bromide
iodine	iod-	iodide

To form the name of the compound, we place the two parts of the name one after another. Therefore, the name of HCl is hydrogen chloride. However, to name PCl_5, we need a way to specify the number of Cl atoms bound to the phosphorus in the molecule. This is done using the following Greek prefixes:

mono- = 1 (often omitted)	hexa- = 6
di- = 2	hepta- = 7
tri- = 3	octa- = 8
tetra- = 4	nona- = 9
penta- = 5	deca- = 10

To name PCl_5, therefore, we add the prefix *penta-* to chloride to give the name phosphorus pentachloride. Notice how easily this allows us to translate the name back into the formula.



TOOLS *Naming binary molecular compounds*

The prefix *mono-* is used when we want to emphasize that only one atom of a particular element is present. For instance, carbon forms two compounds with oxygen, CO and CO_2. To clearly distinguish between them, the first is called carbon monoxide (one of the o's is omitted to make the name easier to pronounce) and the second is carbon dioxide.

As indicated above, the prefix *mono-* is often omitted from a name. Therefore, in general, if there is no prefix before the name of an element, we take it to mean there is only one atom of that element in the molecule. An exception to this is in the names of binary compounds of nonmetals with hydrogen. An example is hydrogen sulfide. The name tells us the compound contains the two elements hydrogen and sulfur. We don't have to be told how many hydrogens are in the molecule because, as you learned earlier, we can use the periodic table to determine the number of hydrogen atoms in molecules of the simple nonmetal hydrides. Sulfur is in Group VIA, so to get to the noble gas column we have to move two steps to the right; the number of hydrogens combined with the atom of sulfur is two. The formula for hydrogen sulfide is therefore H_2S.

EXAMPLE 2.8
Naming Compounds and Writing Formulas

(1) What is the name of $AsCl_3$? (2) What is the formula for dinitrogen tetraoxide?

ANALYSIS: (1) In naming compounds, the first step is to determine what type of compound is involved. Looking at the periodic table, we see that $AsCl_3$ is made up of two nonmetals, so we conclude that it is a molecular compound. Once we've done this, we apply the tool for naming molecular compounds described above.

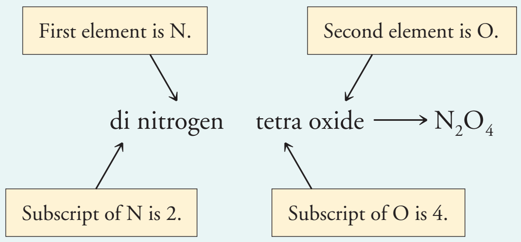
(2) To write the formula from the name, we convert the prefixes to numbers and apply them as subscripts to the chemical symbols of the elements.

After we've discussed ionic compounds this first step in the analysis will be particularly important, because different rules apply depending on the type of compound being named.

SOLUTION: (1) In $AsCl_3$, As is the symbol for arsenic and, of course, Cl is the symbol for chlorine. The first word in the name is just arsenic and the second will contain chloride with an appropriate prefix to indicate number. There are three Cl atoms, so the prefix is tri-. Therefore, the name of the compound is arsenic trichloride.

(2) As we did earlier for phosphorus pentachloride, we convert the prefixes to numbers and apply them as subscripts.

Sometimes we drop the *a* before an *o* for ease of pronunciation. N_2O_4 would then be named dinitrogen tetroxide.



ARE THE ANSWERS REASONABLE? To feel comfortable with the answers, be sure to double-check for careless errors. Next, take your answers and reverse the process. Does arsenic trichloride result in the original formula, $AsCl_3$? Does N_2O_4 have a name of dinitrogen tetraoxide? We can say yes to both and have confidence in our work.

Practice Exercise 25: Name the following compounds using Greek prefixes when needed: (a) PCl_3, (b) SO_2, and (c) Cl_2O_7. (Hint: See the list of prefixes above.)

Practice Exercise 26: Write formulas for the following compounds: (a) arsenic pentachloride, (b) sulfur hexachloride, and (c) disulfur dichloride.

Common names exist for many molecular compounds

Not every compound is named according to the systematic procedure described above. Many familiar substances were discovered long before a systematic method for naming them had been developed, and they acquired common names that are so well known that no attempt has been made to rename them. For example, following the scheme described above we might expect that H_2O would have the name hydrogen oxide (or even dihydrogen monoxide). Although this isn't wrong, the common name water is so well known that it is always used. Another example is ammonia, NH_3, whose odor you have no doubt experienced while using household ammonia solutions or the glass cleaner Windex. Common names are used for the other hydrides of the nonmetals in Group VA as well. The compound PH_3 is called phosphine and AsH_3 is called arsine.

Common names are also used for very complex substances. An example is sucrose, which is the chemical name for table sugar, $C_{12}H_{22}O_{11}$. The structure of this compound is pretty complex, and its name assigned following the systematic method is equally complex. It is much easier to say the simple name sucrose, and be understood, than to struggle with the cumbersome systematic name for this common compound.

For ionic compounds, the name of the cation precedes that of the anion

TOOLS
Naming ionic compounds

In naming ionic compounds, our goal is the same as in naming molecular substances—we want a name that someone else could use to reconstruct the formula. The system we use here, however, is somewhat different than for molecular compounds.

For ionic compounds, the name of the cation is given first, followed by the name of the anion. This is the same as the sequence in which the ions appear in the formula. If the metal in the compound forms only one cation, such as Na^+ or Ca^{2+}, the cation is specified by just giving the English name of the metal. The anion in a binary compound is formed from a nonmetal and its name is created by adding the suffix *-ide* to the stem of the name for the nonmetal just as we did for molecular compounds. An example is KBr, potassium bromide. Table 2.9 lists some common **monatomic** (one-atom) negative ions and their names. It is also useful to know that the *-ide* suffix is usually used only for monatomic ions, with just two common exceptions—*hydroxide ion* (OH^-) and *cyanide ion* (CN^-).[4]

[4] If the name of a compound ends in *-ide* and it isn't either a hydroxide or a cyanide, you can feel confident the substance is a binary compound.

TABLE 2.9 Monatomic Negative Ions

H^-	hydride	N^{3-}	nitride	O^{2-}	oxide	F^-	fluoride
C^{4-}	carbide	P^{3-}	phosphide	S^{2-}	sulfide	Cl^-	chloride
Si^{4-}	silicide	As^{3-}	arsenide	Se^{2-}	selenide	Br^-	bromide
				Te^{2-}	telluride	I^-	iodide

To form the name of an ionic compound, we simply specify the names of the cation and anion. *Use of prefixes to identify the number of cations and anions would be redundant and therefore prefixes are never used when naming ionic compounds.* The reason is that once we know what the ions are, we can assemble the formula correctly just by taking them in a ratio that gives electrical neutrality.

■ To keep the name as simple as possible, we give the minimum amount of information necessary to be able to reconstruct the formula. To write the formula of an ionic compound, we only need the formulas of the ions.

EXAMPLE 2.9
Naming Compounds and Writing Formulas

(a) What is the name of $SrBr_2$? (b) What is the formula for aluminum selenide?

ANALYSIS: The first step in naming a compound is to determine the type of compound it is. As you've already seen, there are slightly different rules for ionic and molecular substances. What clues do we have there that these are ionic?

(a) The element Sr is a metal and Br is a nonmetal. Compounds of a metal and nonmetal are ionic, so we use the rules for naming ionic compounds.

(b) Aluminum is a metal. The only compounds of metals that we've discussed are ionic, so we'll proceed on that assumption. The *-ide* ending of selenide suggests the anion is composed of a single atom of a nonmetal. The only one that begins with the letters "selen-" is selenium, Se. (See the table inside the front cover.)

SOLUTION: (a) The compound is composed of the ions Sr^{2+} (an element from Group IIA) and Br^- (an element from Group VIIA). The cation simply takes the name of the metal, which is strontium. The anion's name is derived from bromine by replacing *-ine* with *-ide*; it is the bromide ion. The name of the compound is strontium bromide.

(b) The name tells us that the cation is the aluminum ion, Al^{3+}. The anion is formed from selenium (Group VIA), and its formula is Se^{2-}. The correct formula must represent an electrically neutral formula unit. Using the number of charges on one ion as the subscript of the other, the formula is Al_2Se_3.

ARE THE ANSWERS REASONABLE? First, we review the analysis and check to be sure we've applied the correct rules, which we have. Next we can reverse the process to be sure our name strontium bromide does mean $SrBr_2$, and that it is reasonable to call Al_2Se_3 aluminum selenide.

Practice Exercise 27: Give the correct formulas for (a) potassium oxide, (b) barium bromide, (c) sodium nitride, and (d) aluminum sulfide. (Hint: Recall what the ending *-ide* means.)

Practice Exercise 28: Give the correct names for (a) $AlCl_3$, (b) BaS, (c) NaBr, and (d) CaF_2.

When a metal can form more than one ion, the charge is indicated with a Roman numeral

Many of the transition metals and post-transition metals are able to form more than one positive ion. Iron, a typical example, forms ions with either a 2+ or a 3+ charge (Fe^{2+} or Fe^{3+}). Compounds that contain these different iron ions have different formulas, so in their names it is necessary to specify which iron ion is present.

FACETS OF CHEMISTRY 2.3

Another System for Naming Ionic Compounds

The Stock system, which we use when naming compounds of metals able to form more than one positive ion, is a relatively recent development in inorganic nomenclature. A slightly different method existed before it.

In the older system, the suffix *-ous* is used to specify the ion with the lower charge and the suffix *-ic* is used to specify the ion with the higher charge. With this method, we also use the Latin stem for elements whose symbols are derived from their Latin names. Some examples are

Fe^{2+}	ferrous ion	$FeCl_2$	ferrous chloride
Fe^{3+}	ferric ion	$FeCl_3$	ferric chloride
Cu^{+}	cuprous ion	$CuCl$	cuprous chloride
Cu^{2+}	cupric ion	$CuCl_2$	cupric chloride

One difficulty with this method is that it does not specify what the charges on the metal ions are, so it becomes necessary to memorize the ions formed by the metals. Additional examples are given in the table below. Notice that mercury is an exception; we use the English stem when naming its ions.

Even though the Stock system is now preferred, some chemical companies still label bottles of chemicals using the old system. These old names also appear in the older scientific literature, which still holds much excellent data. This means that you may need to learn both systems.

The older system of nomenclature is still found on the labels of many laboratory chemicals. This bottle contains copper(II) sulfate, which according to the older system of nomenclature is called cupric sulfate. *(Michael Watson.)*

Cr^{2+}	chromous	Mn^{2+}	manganous	Fe^{2+}	ferrous	Co^{2+}	cobaltous
Cr^{3+}	chromic	Mn^{3+}	manganic	Fe^{3+}	ferric	Co^{3+}	cobaltic
Au^{+}	aurous	Hg_2^{2+}	mercurous	Sn^{2+}	stannous	Pb^{2+}	plumbous
Au^{3+}	auric	Hg^{2+}	mercuric	Sn^{4+}	stannic	Pb^{4+}	plumbic

Originally the cations with different charges were given names that distinguished the higher charge from the lower one. That system is described in Facets of Chemistry 2.3 for those who are interested.

Alfred Stock (1876–1946), a German inorganic chemist, was one of the first scientists to warn the public of the dangers of mercury poisoning.

The currently preferred method for naming ions of metals that can have more than one charge in compounds is called the **Stock system.** Here we use the English name followed, *without a space*, by the numerical value of the charge written as a Roman numeral in parentheses.[5] Examples of using the Stock system are shown below.

Fe^{2+}	iron(II)	$FeCl_2$	iron(II) chloride
Fe^{3+}	iron(III)	$FeCl_3$	iron(III) chloride
Cr^{2+}	chromium(II)	CrS	chromium(II) sulfide
Cr^{3+}	chromium(III)	Cr_2S_3	chromium(III) sulfide

TOOLS
Using the Stock system

Remember that *the Roman numeral equals the positive charge on the metal ion*; it is not necessarily a subscript in the formula. For example, copper forms two oxides, one containing the Cu^{+} ion and the other containing the Cu^{2+} ion. Their formulas are Cu_2O and CuO and their names are as follows[6]:

[5] Silver and nickel are almost always found in compounds as Ag^{+} and Ni^{2+}, respectively. Therefore, AgCl and $NiCl_2$ are almost always called simply silver chloride and nickel chloride.

[6] For some metals, such as copper and lead, one of their ions is much more commonly found in compounds than any of their others. For example, most common copper compounds contain Cu^{2+} and most common lead compounds contain Pb^{2+}. For compounds of these metals, if the charge is not indicated by a Roman numeral, we assume the ion present has a 2+ charge. Thus, it not unusual to find $PbCl_2$ called lead chloride, or for $CuCl_2$ to be called copper chloride.

Cu^+	copper(I)	Cu_2O	copper(I) oxide
Cu^{2+}	copper(II)	CuO	copper(II) oxide

These copper compounds illustrate that in deriving the formula from the name, you must figure out the formula from the ionic charges, as discussed in Section 2.8 and illustrated in the preceding example.

EXAMPLE 2.10
Naming Compounds and Writing Formulas

The compound $MnCl_2$ has a number of commercial uses, including disinfecting, the manufacture of batteries, and purifying natural gas. What is the name of the compound?

ANALYSIS: The first step is to determine the kind of compound involved so we can apply the appropriate rules. The compound here is made up of a metal and a nonmetal, so we use the rules for naming ionic compounds.

To name the cation, we need to know whether the metal is one that forms more than one positive ion. Manganese (Mn) is a transition element, and transition elements often do form more than one cation, so we should apply the Stock method as our tool. To do this, we need to determine the charge on the manganese cation. We can figure this out because the sum of the charges on the Mn and Cl ions must equal zero, and because the only ion chlorine forms has a single negative charge.

SOLUTION: The anion of chlorine (the chloride ion) is Cl^-, so a total of two negative charges are supplied by the two Cl^- ions. Therefore, for $MnCl_2$ to be electrically neutral, the Mn ion must carry two positive charges, 2+. The cation is named as manganese(II), and the name of the compound is manganese(II) chloride.

IS THE ANSWER REASONABLE? Performing a quick check of the arithmetic assures us we've got the correct charges on the ions. Everything appears to be okay.

EXAMPLE 2.11
Naming Compounds and Writing Formulas

What is the formula for cobalt(III) fluoride?

ANALYSIS: To answer this question, we first need to determine the charges on the two ions using the tool above. Then we assemble them into a chemical formula being sure to achieve an electrically neutral formula unit.

SOLUTION: Cobalt(III) corresponds to Co^{3+}. The fluoride ion is F^-. To obtain an electrically neutral substance, we must have three F^- ions for each Co^{3+} ion, so the formula is CoF_3.

IS THE ANSWER REASONABLE? We can check to see that we have the correct formulas of the ions and that we've combined them to achieve an electrically neutral formula unit. This will tell us we've obtained the correct answer.

Practice Exercise 29: Name the compounds (a) K_2S, (b) Mg_3P_2, (c) $NiCl_2$, and (d) Fe_2O_3. Use the Stock system where appropriate. (Hint: Determine which metals can have more than one charge.)

Practice Exercise 30: Write formulas for (a) aluminum sulfide, (b) strontium fluoride, (c) titanium(IV) oxide, and (d) gold(III) oxide.

Similar rules apply to naming ionic compounds that contain polyatomic ions

TOOLS *Naming with polyatomic ions*

The extension of the nomenclature system to include ionic compounds containing polyatomic ions is straightforward. Most of the polyatomic ions listed in Table 2.8 are anions and their names are used as the second word in the name of the compound. For example, Na_2SO_4 contains the sulfate ion, SO_4^{2-}, and is called sodium sulfate. Similarly, $Cr(NO_3)_3$ contains the nitrate ion, NO_3^-. Chromium is a transition element, and in this compound its charge must be 3+ to balance the negative charges of three NO_3^- ions. Therefore, $Cr(NO_3)_3$ is called chromium(III) nitrate.

It is important that you learn the formulas (including charges) and the names of the polyatomic ions in Table 2.8. You will encounter them frequently throughout your chemistry course.

Among the ions in Table 2.8, the only cation that forms compounds which can be isolated is ammonium ion, NH_4^+. It forms ionic compounds such as NH_4Cl (ammonium chloride) and $(NH_4)_2SO_4$ (ammonium sulfate), even though NH_4^+ is not a metal cation. Notice that the latter compound is composed of two polyatomic ions.

EXAMPLE 2.12
Naming Compounds and Writing Formulas

What is the name of $Mg(ClO_4)_2$, a compound used commercially for removing moisture from gases?

ANALYSIS: To answer this question, it is essential that you recognize that the compound contains a polyatomic ion. If you've learned the contents of Table 2.8, you know that "ClO_4" is the formula (without the charge) of the perchlorate ion. With the charge, the ion's formula is ClO_4^-. We also have to decide whether we need to apply the Stock system in naming the metal Mg. We use the naming tool for polyatomic ion compounds to complete the exercise.

SOLUTION: Magnesium is in Group IIA, and forms only the ion Mg^{2+}. Therefore, we don't need to use the Stock system in naming the cation; it is named simply as "magnesium." The anion is perchlorate, so the name of $Mg(ClO_4)_2$ is magnesium perchlorate.

IS THE ANSWER REASONABLE? We can check to be sure we've named the anion correctly, and we have. The metal is magnesium, which only forms Mg^{2+}. Therefore, the answer seems to be correct.

Practice Exercise 31: What are the names of (a) Li_2CO_3 and (b) $Fe(OH)_3$? (Hint: Recall the names of the polyatomic ions and the positions of Li and Fe in the periodic table.)

Practice Exercise 32: Write the formulas for (a) potassium chlorate and (b) nickel(II) phosphate.

Hydrates are named using Greek prefixes

Earlier we discussed compounds called hydrates, such as $CuSO_4 \cdot 5H_2O$. Usually, hydrates are ionic compounds whose crystals contain water molecules in fixed proportions relative to the ionic substance. To name them, we provide two pieces of information: the name of the ionic compound and the number of water molecules in the formula. The number of water molecules is specified using the Greek prefixes mentioned earlier (mono-, di-, tri-, etc.), which precede the word "hydrate." Thus, $CuSO_4 \cdot 5H_2O$ is named as "copper sulfate *pentahydrate*." Similarly, $CaSO_4 \cdot 2H_2O$ is named calcium sulfate dihydrate, and $FeCl_3 \cdot 6H_2O$ is iron(III) chloride hexahydrate.[7]

[7] Chemical suppliers (who do not always follow current rules of nomenclature) sometimes indicate the number of water molecules using a number and a dash. For example, one supplier lists $Ca(NO_3)_2 \cdot 4H_2O$ as "Calcium nitrate, 4-hydrate."

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Naming a compound requires that we select the rules that apply

In this chapter we've discussed how to name two classes of compounds, molecular and ionic, and you saw that slightly different rules apply to each. To name chemical compounds successfully we need to make a series of decisions based on the rules we just covered. At this point we can summarize this decision process in a flowchart such as the one shown in Figure 2.28. The example below illustrates how to use the flowchart, and when working on the Review Problems, you may want to refer to Figure 2.28 until you are able to develop the skills that will enable you to work without it.

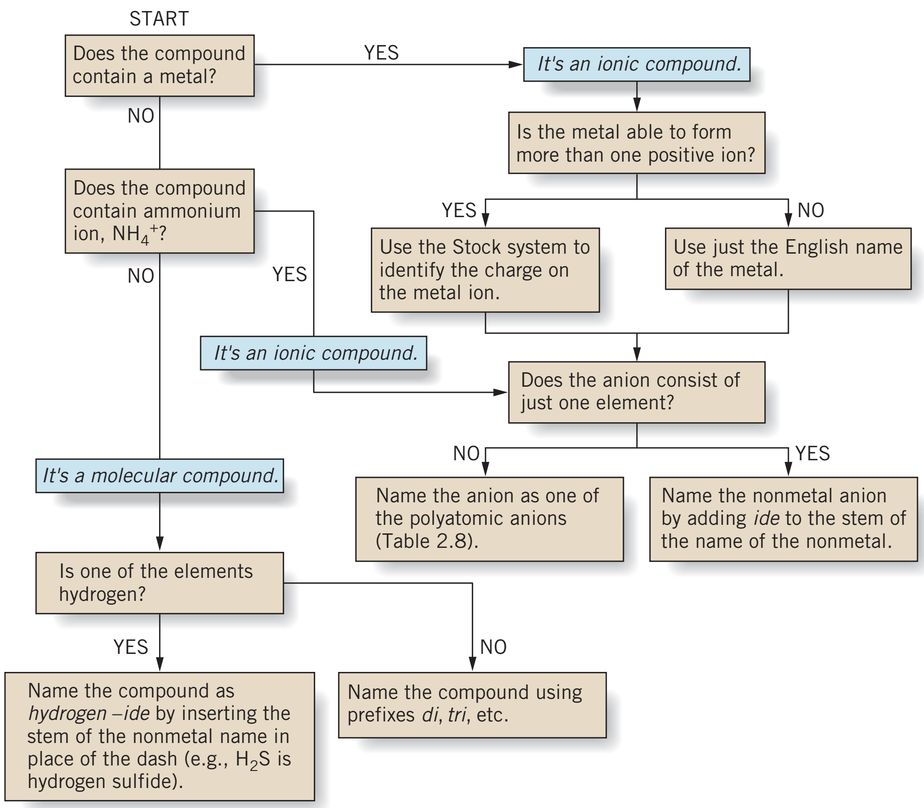

FIG. 2.28 Flowchart for naming molecular and ionic compounds.

EXAMPLE 2.13
Applying the Rules for Naming Compounds

What is the name of (a) $CrCl_3$, (b) P_4S_3, and (c) NH_4NO_3?

ANALYSIS: For each compound, we use the tools summarized in Figure 2.28 and proceed through the decision processes to arrive at the way to name the compound. As you read through the solution below, be sure to refer to Figure 2.28 so you can see how we decide which rules we need to apply.

SOLUTION: (a) Starting at the top of Figure 2.28, we first determine that the compound contains a metal (Cr), so it's an ionic compound. Next, we see that the metal is a transition element, and chromium is one of those that forms more than one cation, so we have to apply the Stock method. To do this, we need to know what the charge is on the metal ion. We can figure this out using the charge on the anion and the fact that the compound must be electrically neutral overall. The anion is formed from chlorine, so its charge must be 1− (the anion is Cl^-). Therefore, the metal ion must be Cr^{3+}; we name the metal as *chromium(III)*. Next, we see that there is only one nonmetallic element in the compound, Cl, so the name of the anion ends in *-ide*; it's the *chloride* ion. The compound $CrCl_3$ is therefore named *chromium(III) chloride*.

(b) Once again, we start at the top. First we determine that the compound doesn't contain a metal. It also doesn't contain NH_4, so the compound is molecular. It doesn't contain hydrogen, so we are led to the decision that we must use Greek prefixes to specify the numbers of atoms of each element. Applying the procedure on page 71, the name of the compound P_4S_3 is *tetraphosphorus trisulfide*.

(c) We begin at the top. Studying the formula, we see that it does not contain the symbol for a metal, so we proceed down the left side of the figure. The formula does contain NH_4 which indicates the compound contains the *ammonium* ion, NH_4^+ (it's an ionic compound). The rest of the formula is NO_3, which consists of more than one atom. This suggests the polyatomic anion, NO_3^- (*nitrate* ion). The name of the compound NH_4NO_3 is *ammonium nitrate*.

ARE THE ANSWERS REASONABLE? To check the answers in a problem of this kind, review the decision processes which led you to the names. In part (a), you can check to be sure you've calculated the charge on the chromium ion correctly. Also, check to be sure you've used the correct names of any polyatomic ions. Doing these things will show we've named the compounds correctly.

Practice Exercise 33: The compound I_2O_5 is used in respirators where it serves to react with highly toxic carbon monoxide to give the much less toxic gas, carbon dioxide. What is the name of I_2O_5? (Hint: Is this a molecular or ionic compound?)

Practice Exercise 34: The compound $Cr(C_2H_3O_2)_3$ is used in the tanning of leather. What is the name for this compound?

SUMMARY

Elements and Atoms. When accurate masses of all the reactants and products in a reaction are measured and compared, no observable changes in mass accompany chemical reactions (the **law of conservation of mass**). The mass ratios of the elements in any compound are constant regardless of the source of the compound or how it is prepared (the **law of definite proportions**). **Dalton's atomic theory** explained the laws of chemical combination by proposing that matter consists of indestructible atoms with masses that do not change during chemical reactions. During a chemical reaction, atoms may change partners, but they are neither created nor destroyed. After Dalton had proposed his theory, it was discovered that whenever two elements form more than one compound, the different masses of one element that combine with a fixed mass of the other are in a ratio of small whole numbers (the **law of multiple proportions**). Using modern instruments such as the **scanning tunneling microscope,** scientists are able to "see" atoms on the surfaces of solids.

Atomic Mass. An element's **atomic mass** (**atomic weight**) is the relative mass of its atoms on a scale in which atoms of carbon-12 have a mass of exactly 12 u (**atomic mass units**). Most elements occur in nature as uniform mixtures of a small number of **isotopes,** whose masses differ slightly. However, all isotopes of an element have very nearly identical chemical properties and the percentages of the isotopes that make up an element are generally so constant throughout the world that we can say that the average mass of their atoms is a constant.

Atomic Structure. Atoms can be split into **subatomic particles,** such as **electrons, protons,** and **neutrons. Nucleons** are particles that make up the atomic **nucleus** and include the protons, each of which carries a single unit of **positive charge** (charge = 1+), and neutrons (no charge). The number of protons is called the **atomic number** (*Z*) of the element. Each element has a different atomic number. The electrons, each with a unit of **negative charge** (charge = 1−) are found outside the nucleus; their number equals the atomic number in a neutral atom. Isotopes of an element have identical atomic numbers but different numbers of neutrons. In more modern terms, an **element** can be defined as a substance whose atoms all have the same number of protons in their nuclei.

The Periodic Table. The search for similarities and differences among the properties of the elements led Mendeleev to discover that when the elements are placed in (approximate) order of increasing atomic mass, similar properties recur at regular, repeating intervals. In the modern **periodic table** the elements are arranged in rows, called **periods,** in order of increasing atomic number. The rows are stacked so that elements in the columns, called **groups** or **families,** have similar chemical and physical properties. The A-group elements (IUPAC Groups 1, 2, and 13–18) are called **representative elements;** the B-group elements (IUPAC Groups 3–12) are called **transition elements.** The two long rows of **inner transition elements** located below the main body of the table consist of the **lanthanides,** which follow La ($Z = 57$), and the **actinides,** which follow Ac ($Z = 89$). Certain groups are given family names: Group IA (Group 1), except for hydrogen, are the **alkali metals** (the alkalis); Group IIA (Group 2), the **alkaline earth metals;** Group VIIA (Group 17), the **halogens;** Group VIIIA (Group 18), the **noble gases**.

Metals, Nonmetals, and Metalloids. Most elements are **metals;** they occupy the lower left-hand region of the periodic table (to the left of a line drawn approximately from boron, B, to astatine, At). **Nonmetals** are found in the upper right-hand region of the table. **Metalloids** occupy a narrow band between the metals and nonmetals.

Metals exhibit a **metallic luster,** tend to be **ductile** and **malleable,** and conduct electricity. Nonmetals tend to be brittle, lack metallic luster, and are nonconductors of electricity. Many nonmetals are gases. Bromine (a nonmetal) and mercury (a metal) are the two elements that are liquids at ordinary room temperature. Metalloids have properties intermediate between those of metals and nonmetals and are **semiconductors** of electricity.

Reactions of Elements to Form Compounds. Nearly every element has the ability to form compounds, although not all combinations are possible. When a chemical reaction takes place, the properties of the substances present at the start disappear and are replaced by the properties of the substances formed in the reaction. Chemical symbols are used to write **chemical formulas,** both for **free elements** that occur as molecules (e.g., **diatomic molecules** of elements such as H_2, O_2, N_2, and Cl_2) and for chemical compounds. **Subscripts** are used to specify how many atoms of each element are present. Some compounds form solids called **hydrates,** which contain water molecules in definite proportions. Heating a hydrate usually drives off the water.

Chemical Equations. A **chemical equation** presents a before-and-after description of a chemical reaction. When **balanced,** an equation contains **coefficients** that make the number of atoms of each kind the same among the **reactants** and the **products.** In this way, the equation conforms to the law of conservation of mass. The physical states of the reactants and products can be specified in an equation by placing the following symbols within parentheses following the chemical formulas: *s*, *l*, *g*, and *aq*, which stand for solid, liquid, gas, and aqueous solution (dissolved in water), respectively.

Molecules and Molecular Compounds. **Molecules** are electrically neutral particles consisting of two or more atoms. The erratic movements of microscopic particles suspended in a liquid (**Brownian motion**) can be interpreted to be caused by collisions with molecules of the liquid. Molecules are held together by **chemical bonds** that arise from the sharing of electrons between atoms. Formulas we write for molecules are **molecular formulas.** Molecular compounds are formed when nonmetals combine with each other. The simple nonmetal hydrides have formulas that can be remembered by the position of the nonmetal in the periodic table. **Organic compounds** are **hydrocarbons,** or compounds considered to be derived from hydrocarbons by replacing H atoms with other atoms.

Ions and Ionic Compounds. **Binary ionic** compounds are formed when metals react with nonmetals. In the reaction, electrons are transferred from a metal to a nonmetal. The metal atom becomes a positive ion (a **cation**); the nonmetal atom becomes a negative ion (an **anion**). The formula of an ionic compound specifies the smallest whole-number ratio of the ions. The smallest unit of an ionic compound is called a **formula unit,** which specifies the smallest whole-number ratio of the ions that produces electrical neutrality. Many ionic compounds also contain **polyatomic ions**—ions that are composed of two or more atoms.

Naming Molecular Compounds. The system of **nomenclature** for **binary** molecular **inorganic compounds** uses a set of **Greek prefixes** to specify the numbers of atoms of each kind in the formula of the compound. The first element in the formula is specified by its English name; the second element takes the suffix *-ide,* which is added to the stem of the English name. For simple nonmetal hydrides, it is not necessary to specify the number of hydrogens in the formula. Many familiar substances as well as very complex molecules are usually identified by **common names.**

Naming Ionic Compounds. In naming an ionic compound, the cation is specified first, followed by the anion. Metal cations take the English name of the element, and when more than one positive ion can be formed by the metal, the **Stock system** is used to identify the amount of positive charge on the cation. This is done by placing a Roman numeral equal to the positive charge in parentheses following the name of the metal. Simple **monatomic** anions are formed by nonmetals and their names are formed by adding the suffix *-ide* to the stem of the nonmetal's name. Only two common polyatomic anions (cyanide and hydroxide) end in the suffix *-ide.*

Properties of Molecular and Ionic Compounds. We found that it is often possible to distinguish ionic compounds from molecular compounds by their ability to **conduct electricity.** Molecular compounds are generally poor electrical conductors, whereas ionic compounds when melted into the liquid state or dissolved in water will conduct electricity readily.

TOOLS FOR PROBLEM SOLVING

We have learned the following concepts which can be applied as tools in solving problems. Study each one carefully so that you know what each is used for. When faced with solving a problem, recall what each tool does and consider whether it will be helpful in finding a solution. This will aid you in selecting the tools you need. If necessary, refer to these tools when working on the exercises in the chapter and the review questions and problems that follow. Remember that tools from Chapter 1 may be needed at times to solve problems in this chapter.

Law of definite proportions *(page 36)* If we know the **mass ratio** of the elements in one sample of a compound, we know the ratio will be the same in a different sample of the same compound.

Law of conservation of mass *(page 36)* The total mass of chemicals present before a reaction starts equals the total mass after the reaction is finished. We can use this law to check whether we have accounted for all the substances formed in a reaction.

Subatomic particles *(page 43* There are three important relationships between the numbers of **protons, neutrons (nucleons)**, and **electrons.** These relationships are

For a neutral atom: number of electrons = number of protons
Atomic number (Z) = number of protons
Mass number (A) = number of protons + number of neutrons

Relative atomic masses *(page 45)* Atomic masses are relative to the mass of a ^{12}C atom that has a mass of exactly 12 atomic mass units (u). Therefore the atomic mass of ^{12}C is exactly 12 u.

Chemical formula *(page 54)* **Subscripts** in a formula specify the number of atoms of each element in one formula unit of the substance. This gives us *atom ratios* that we will find useful when we deal with the compositions of compounds in Chapter 3.

Rules for writing formulas of ionic compounds *(page 66)* The rules permit us to write correct chemical formulas for ionic compounds. You will need to learn to use the periodic table (see below) to remember the charges on the cations and anions of the representative metals and nonmetals.

Polyatomic ions *(page 69)* Certain groups of atoms arrange themselves into stable configurations that we call polyatomic ions. It is very important that you commit to memory the names, formulas, and charges of these ions (which are given in Table 2.8).

Monatomic anion names *(page 73)* The list on this page gives the common names of anions that must be remembered.

Greek prefixes *(page 71)* This page has a list of the Greek prefixes from one to ten that you should know for naming molecular compounds and hydrates.

Rules for naming molecular compounds *(page 71)* These rules give us a logical system for naming binary molecular compounds by specifying the number of each type of atom using Greek prefixes.

Rules for naming ionic compounds *(page 72)* These rules give us a systematic method for naming ionic compounds. The name of the cation is combined with the name of a monatomic anion as given on page 73. These rules are used with slight modification for cations that can have more than one possible charge (see using the **Stock system** below) and for situations where a polyatomic ion is involved (see naming with **polyatomic ions** below)

Using the Stock system *(page 74)* The Stock system specifies the charge of a cation by placing a Roman numeral in parentheses just after the name of the cation. The Stock system and its Roman numerals are only used for cations that can have more than one possible charge.

Naming with polyatomic ions *(page 76)* Naming compounds that contain polyatomic anions is done by specifying the cation name, using the Stock system if needed, and then specifying the polyatomic anion name as given in Table 2.8. The one polyatomic cation, the ammonium ion (NH_4^+), uses its name and then the appropriate name of the anion.

Periodic table The periodic table has several tool icons in this chapter illustrating its use in a variety of different ways. The periodic table is used to find group numbers and period numbers of the elements *(page 49)*. The table will help us recall group names and we can obtain atomic numbers and average masses of the elements *(page 50)*. From an element's position in the periodic table, we can tell whether it's a metal, nonmetal, or metalloid *(page 51)*. From a nonmetal's position in the periodic table we can write the formula of its simple hydride *(page 60)* and predict the charge of monatomic anions *(page 65)*. For the metals in Groups IA and IIA, we can use the elements' positions in the periodic table to obtain the charges on their ions *(page 65)*.

QUESTIONS, PROBLEMS, AND EXERCISES

Answers to problems whose numbers are printed in color are given in Appendix B. More challenging problems are marked with asterisks. ILW = Interactive Learningware solution is available at www.wiley.com/college/brady. OH = an Office Hours video is available for this problem.

REVIEW QUESTIONS

Laws of Chemical Combination and Dalton's Theory

2.1 Name and state the two laws of chemical combination discussed in this chapter.

2.2 Why didn't the existence of isotopes affect the apparent validity of the atomic theory?

2.3 In your own words, describe how Dalton's theory explains the law of conservation of mass and the law of definite proportions.

2.4 Which of the laws of chemical combination is used to define the term *compound*?

2.5 Describe what you need to do in the laboratory to test (a) the law of conservation of mass, (b) the law of definite proportions, and (c) the law of multiple proportions.

Atomic Masses and Atomic Structure

2.6 What are the names, symbols, and electrical charges of the three subatomic particles introduced in this chapter?

2.7 Where in an atom is nearly all of its mass concentrated? Explain your answer in terms of the particles that contribute to this mass.

2.8 What is a *nucleon*? Which ones have we studied?

2.9 Define the terms *atomic number* and *mass number.* What symbols are used to designate these terms?

2.10 Consider the symbol ${}^{a}_{b}X$, where X stands for the chemical symbol for an element. What information is given in locations (a) a and (b) b?

2.11 Write the symbols of the isotopes that contain the following. (Use the table of atomic masses and numbers printed inside the front cover for additional information, as needed.)
(a) An isotope of iodine whose atoms have 78 neutrons.
(b) An isotope of strontium whose atoms have 52 neutrons.
(c) An isotope of cesium whose atoms have 82 neutrons.
(d) An isotope of fluorine whose atoms have 9 neutrons.

The Periodic Table

2.12 In the compounds formed by Li, Na, K, Rb, and Cs with chlorine, how many atoms of Cl are there per atom of the metal? In the compounds formed by Be, Mg, Ca, Sr, and Ba with chlorine, how many atoms of Cl are there per atom of metal? How did this kind of information lead Mendeleev to develop his periodic table?

2.13 On what basis did Mendeleev construct his periodic table? On what basis are the elements arranged in the modern periodic table?

2.14 On the basis of their positions in the periodic table, why is it not surprising that strontium-90, a dangerous radioactive isotope of strontium, replaces calcium in newly formed bones?

2.15 In the refining of copper, sizable amounts of silver and gold are recovered. Why is this not surprising?

2.16 Why would you reasonably expect cadmium to be a contaminant in zinc but not in silver?

2.17 Using the symbol for nitrogen, N, indicate what information is conveyed by the superscripts before and after the symbol and by subscripts before and after the symbol.

2.18 Make a rough sketch of the periodic table and mark off those areas where you would find (a) the representative elements, (b) the transition elements, and (c) the inner transition elements.

2.19 Which of the following is
(a) an alkali metal: Ca, Cu, In, Li, S?
(b) a halogen: Ce, Hg, Si, O, I?
(c) a transition element: Pb, W, Ca, Cs, P?
(d) a noble gas: Xe, Se, H, Sr, Zr?
(e) a lanthanide element: Th, Sm, Ba, F, Sb?
(f) an actinide element: Ho, Mn, Pu, At, Na?
(g) an alkaline earth metal: Mg, Fe, K, Cl, Ni?

Physical Properties of Metals, Nonmetals, and Metalloids

2.20 Name five physical properties that we usually observe for metals.

2.21 Why is mercury used in thermometers? Why is tungsten used in lightbulbs?

2.22 Which nonmetals occur as monatomic gases (gases whose particles consist of single atoms)?

2.23 Which two elements exist as liquids at room temperature and pressure?

2.24 Which physical property of metalloids distinguishes them from metals and nonmetals?

2.25 Sketch the shape of the periodic table and mark off those areas where we find (a) metals, (b) nonmetals, and (c) metalloids.

2.26 Most periodic tables have a heavy line that looks like a staircase starting from boron down to polonium. What information does this line convey?

2.27 Which metals can you think of that are commonly used to make jewelry? Why isn't iron used to make jewelry? Why isn't potassium used?

2.28 What trends (regular changes in physical or chemical properties) in the periodic table have been mentioned in this chapter?

2.29 Find a periodic table on the Internet that lists physical properties of the elements. Can you distinguish trends in the periodic table based on (a) melting point, (b) boiling point, or (c) density?

Chemical Formulas

2.30 What are two ways to interpret a chemical symbol?

2.31 What is the difference between an atom and a molecule?

2.32 Write the formulas and names of the nonmetallic elements that exist in nature as diatomic molecules.

Chemical Equations

2.33 What do we mean when we say a chemical equation is *balanced*? Why do we balance chemical equations?

2.34 For a chemical reaction, what do we mean by the term *reactants*? What do we mean by the term *products*?

2.35 The combustion of a thin wire of magnesium metal (Mg) in an atmosphere of pure oxygen produces the brilliant light of a flashbulb, once commonly used in photography. After the reaction, a thin film of magnesium oxide is seen on the inside of the bulb. The equation for the reaction is

$$2\text{Mg} + \text{O}_2 \longrightarrow 2\text{MgO}$$

(a) State in words how this equation is read.
(b) Give the formula(s) of the reactants.
(c) Give the formula(s) of the products.
(d) Rewrite the equation to show that Mg and MgO are solids and O_2 is a gas.

2.36 The chemical equation for the combustion of octane (C_8H_{18}), a component of gasoline, is

$$2\text{C}_8\text{H}_{18} + 25\text{O}_2 \longrightarrow 16\text{CO}_2 + 18\text{H}_2\text{O}$$

Rewrite the equation so that it specifies octane and water as liquids and oxygen and carbon dioxide (CO_2) as gases.

Molecular Compounds of Nonmetals

2.37 Which are the only elements that exist as free, individual atoms when not chemically combined with other elements?

2.38 Write chemical formulas for molecules of elemental sulfur and phosphorus mentioned in this chapter.

2.39 Which kind of elements normally combine to form molecular compounds?

2.40 Without referring to Table 2.4, but using the periodic table, write chemical formulas for the simplest hydrogen compounds of (a) carbon, (b) nitrogen, (c) tellurium, and (d) iodine.

2.41 The simplest hydrogen compound of phosphorus is phosphine, a highly flammable and poisonous compound with an odor of decaying fish. What is the formula for phosphine?

2.42 Astatine, a radioactive member of the halogen family, forms a compound with hydrogen. Predict its chemical formula.

2.43 Under appropriate conditions, tin can be made to form a simple molecular compound with hydrogen. Predict its formula.

2.44 Write the chemical formulas for (a) methane, (b) ethane, (c) propane, and (d) butane. Give one practical use for each of these hydrocarbons.

2.45 What are the formulas of (a) methanol and (b) ethanol?

2.46 What is the formula for the alkane that has 10 carbon atoms?

2.47 Candle wax is a mixture of hydrocarbons, one of which is an alkane with 23 carbon atoms. What is the formula for this hydrocarbon?

2.48 The formula for a compound is correctly given as $C_6H_{12}O_6$. State two reasons why we expect this to be a molecular compound, rather than an ionic compound.

2.49 Explore the Internet and find a reliable source of structures for molecular compounds. For questions 44 to 47 print out the ball-and-stick and space-filling models of the compounds mentioned.

Ionic Compounds

2.50 Describe what kind of event must occur (involving electrons) if the atoms of two different elements are to react to form (a) an ionic compound or (b) a molecular compound.

2.51 With what kind of elements do metals react?

2.52 With what kind of elements do nonmetals react?

2.53 Why are nonmetals found in more compounds than are metals, even though there are fewer nonmetals than metals?

2.54 What is an ion? How does it differ from an atom or a molecule?

2.55 Why do we use the term *formula unit* for ionic compounds instead of the term *molecule*?

2.56 Most compounds of aluminum are ionic, but a few are molecular. How do we know that Al_2Cl_6 is molecular?

2.57 Consider the sodium atom and the sodium ion.
(a) Write the chemical symbol of each.
(b) Do these particles have the same number of nuclei?
(c) Do they have the same number of protons?
(d) Could they have different numbers of neutrons?
(e) Do they have the same number of electrons?

2.58 Define *cation, anion,* and *polyatomic ion.*

2.59 How many electrons has a titanium atom lost if it has formed the ion Ti^{4+}? What are the total numbers of protons and electrons in a Ti^{4+} ion?

2.60 If an atom gains an electron to become an ion, what kind of electrical charge does the ion have?

2.61 How many electrons has a nitrogen atom gained if it has formed the ion N^{3-}? How many protons and electrons are in an N^{3-} ion?

2.62 What is wrong with the formula $RbCl_3$? What is wrong with the formula SNa_2?

2.63 A student wrote the formula for an ionic compound of titanium as Ti_2O_4. What is wrong with this formula? What should the formula be?

2.64 What are the formulas of the ions formed by (a) iron, (b) cobalt, (c) mercury, (d) chromium, (e) tin, and (f) copper?

2.65 Which of the following formulas are incorrect? (a) NaN_2, (b) RbCl, (c) K_2S, (d) Al_2Cl_3, (e) MgO_2

2.66 What are the formulas (including charges) for (a) cyanide ion, (b) ammonium ion, (c) nitrate ion, (d) sulfite ion, (e) chlorate ion, and (f) sulfate ion?

2.67 What are the formulas (including charges) for (a) hypochlorite ion, (b) bisulfate ion, (c) phosphate ion, (d) dihydrogen phosphate ion, (e) permanganate ion, and (f) oxalate ion?

2.68 What are the names of the following ions? (a) $Cr_2O_7^{2-}$, (b) OH^-, (c) $C_2H_3O_2^-$, (d) CO_3^{2-}, (e) CN^-, (f) ClO_4^-

2.69 From what you have learned in Section 2.8, write correct balanced equations for the reactions between (a) calcium and chlorine, (b) magnesium and oxygen, (c) aluminum and oxygen, and (d) sodium and sulfur.

2.70 Write the balanced equations (including phases) for the reactions described below: (a) Solid iron(III) hydroxide reacts with gaseous hydrogen chloride forming water and iron(III) chloride. (b) Aqueous silver nitrate is reacted with aqueous barium chloride to form solid silver chloride and aqueous barium nitrate.

2.71 Write the balanced equations (including phases) for the reactions described below: (a) Gaseous propane reacts with gaseous oxygen to form carbon dioxide and water. (b) Sodium metal is added to water and the products are aqueous sodium hydroxide and hydrogen gas.

Naming Compounds

2.72 What is the difference between a binary compound and one that is diatomic? Give examples that illustrate this difference.

2.73 In naming the compounds discussed in this chapter, why is it important to know whether a compound is molecular or ionic?

2.74 In naming ionic compounds of the transition elements, why is it essential to know the charge on the anion?

2.75 Describe (a) the three situations in which Greek prefixes are used and (b) when Roman numerals are used.

REVIEW PROBLEMS

Laws of Chemical Combination

OH **2.76** Ammonia is composed of hydrogen and nitrogen in a ratio of 9.33 g of nitrogen to 2.00 g of hydrogen. If a sample of ammonia contains 6.28 g of hydrogen, how many grams of nitrogen does it contain?

2.77 A compound of phosphorus and chlorine used in the manufacture of a flame retardant treatment for fabrics contains 1.20 grams of phosphorus for every 4.12 g of chlorine. Suppose a sample of this compound contains 6.22 g of chlorine. How many grams of phosphorus does it contain?

2.78 Refer to the data about ammonia in Problem 2.76. If 4.56 g of nitrogen combined completely with hydrogen to form ammonia, how many grams of ammonia would be formed?

2.79 Refer to the data about the phosphorus–chlorine compound in Problem 2.77. If 12.5 g of phosphorus combined completely with chlorine to form this compound, how many grams of the compound would be formed?

2.80 Molecules of a certain compound of nitrogen and oxygen contain one atom each of N and O. In this compound there are 1.143 g of oxygen for each 1.000 g of nitrogen. Molecules of a different compound of nitrogen and oxygen contain one atom of N and two atoms of O. How many grams of oxygen would be combined with each 1.000 g of nitrogen in the second compound?

2.81 Tin forms two compounds with chlorine. In one of them (compound 1), there are two Cl atoms for each Sn atom; in the other (compound 2), there are four Cl atoms for each Sn atom. When combined with the same mass of tin, what would be the ratio of the masses of chlorine in the two compounds? In compound 1, 0.597 g of chlorine is combined with each 1.000 g of tin. How many grams of chlorine would be combined with 1.000 g of tin in compound 2?

Atomic Masses and Isotopes

2.82 The chemical substance in natural gas is a compound called methane. Its molecules are composed of carbon and hydrogen, and each molecule contains four atoms of hydrogen and one atom of carbon. In this compound, 0.33597 g of hydrogen is combined with 1.0000 g of carbon-12. Use this information to calculate the atomic mass of the element hydrogen.

OH **2.83** A certain element *X* forms a compound with oxygen in which there are two atoms of *X* for every three atoms of O. In this compound, 1.125 g of *X* is combined with 1.000 g of oxygen. Use the average atomic mass of oxygen to calculate the average atomic mass of *X*. Use your calculated atomic mass to identify the element *X*.

2.84 If an atom of carbon-12 had been assigned a relative mass of 24.0000 u, what would be the average atomic mass of hydrogen relative to this mass?

2.85 One atom of ^{109}Ag has a mass that is 9.0754 times that of a ^{12}C atom. What is the atomic mass of this isotope of silver expressed in atomic mass units?

Atomic Structure

ILW **2.86** Naturally occurring copper is composed of 69.17% of ^{63}Cu, with an atomic mass of 62.9396 u, and 30.83% of ^{65}Cu, with an atomic mass of 64.9278 u. Use these data to calculate the average atomic mass of copper.

2.87 Naturally occurring magnesium (one of the elements in milk of magnesia) is composed of 78.99% of ^{24}Mg (atomic mass, 23.9850 u), 10.00% of ^{25}Mg (atomic mass, 24.9858 u), and 11.01% of ^{26}Mg (atomic mass, 25.9826 u). Use these data to calculate the average atomic mass of magnesium.

ILW **2.88** Give the numbers of neutrons, protons, and electrons in the atoms of each of the following isotopes. (Use the table of atomic masses and numbers printed inside the front cover for additional information, as needed.) (a) radium-226, (b) ^{206}Pb, (c) carbon-14, (d) ^{23}Na

2.89 Give the numbers of electrons, protons, and neutrons in the atoms of each of the following isotopes. (As necessary, consult the table of atomic masses and numbers printed inside the front cover.) (a) cesium-137, (b) ^{238}U, (c) iodine-131, (d) ^{197}Au

Chemical Formulas

2.90 The compound $Cr(C_2H_3O_2)_3$ is used in the tanning of leather. How many atoms of each element are given in this formula?

2.91 Asbestos, a known cancer-causing agent, has as a typical formula, $Ca_3Mg_5(Si_4O_{11})_2(OH)_2$. How many atoms of each element are given in this formula?

2.92 Epsom salts is a hydrate of magnesium sulfate, $MgSO_4 \cdot 7H_2O$. What is the formula of the substance that remains when Epsom salts is completely dehydrated?

2.93 Rochelle salt is the tetrahydrate of $KNaC_4H_4O_6$. Write the formula for Rochelle salt.

2.94 How many atoms of each element are represented in each of the following formulas? (a) $K_2C_2O_4$, (b) H_2SO_3, (c) $C_{12}H_{26}$, (d) $HC_2H_3O_2$, (e) $(NH_4)_2HPO_4$

2.95 How many atoms of each kind are represented in the following formulas? (a) Na_3PO_4, (b) $Ca(H_2PO_4)_2$, (c) C_4H_{10}, (d) $Fe_3(AsO_4)_2$, (e) $C_3H_5(OH)_3$

2.96 How many atoms of each kind are represented in the following formulas? (a) $Ni(ClO_4)_2$, (b) $CuCO_3$, (c) $K_2Cr_2O_7$, (d) CH_3CO_2H, (e) $(NH_4)_2HPO_4$

2.97 How many atoms of each kind are represented in the following formulas? (a) $CH_3CH_2CO_2C_3H_7$, (b) $MgSO_4 \cdot 7H_2O$, (c) $KAl(SO_4)_2 \cdot 12H_2O$, (d) $Cu(NO_3)_2$, (e) $(CH_3)_3COH$

2.98 How many atoms of each element are represented in each of the following expressions? (a) $3N_2O$, (b) $4NaHCO_3$, (c) $2CuSO_4 \cdot 5H_2O$

OH **2.99** How many atoms of each element are represented in each of the following expressions? (a) $7CH_3CO_2H$, (b) $2(NH_2)_2CO$, (c) $5K_2Cr_2O_7$

Chemical Equations

2.100 Consider the balanced equation

$$2Fe(NO_3)_3 + 3Na_2CO_3 \longrightarrow Fe_2(CO_3)_3 + 6NaNO_3$$

(a) How many atoms of Na are on each side of the equation?
(b) How many atoms of C are on each side of the equation?
(c) How many atoms of O are on each side of the equation?

2.101 Consider the balanced equation for the combustion of octane, a component of gasoline:

$$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$$

(a) How many atoms of C are on each side of the equation?
(b) How many atoms of H are on each side of the equation?
(c) How many atoms of O are on each side of the equation?

Ionic Compounds

2.102 Use the periodic table, but not Table 2.6, to write the symbols for the ions of (a) K, (b) Br, (c) Mg, (d) S, and (e) Al.

2.103 Use the periodic table, but not Table 2.6, to write the symbols for ions of (a) barium, (b) oxygen, (c) fluorine, (d) strontium, and (e) rubidium.

OH **2.104** Write formulas for ionic compounds formed between (a) Na and Br, (b) K and I, (c) Ba and O, (d) Mg and Br, and (e) Ba and F.

2.105 Write the formulas for the ionic compounds formed by the following transition metals with the chloride ion, Cl^-: (a) chromium, (b) iron, (c) manganese, (d) copper, and (e) zinc.

2.106 Write formulas for the ionic compounds formed from (a) K^+ and nitrate ion, (b) Ca^{2+} and acetate ion, (c) ammonium ion and Cl^-, (d) Fe^{3+} and carbonate ion, and (e) Mg^{2+} and phosphate ion.

2.107 Write formulas for the ionic compounds formed from (a) Zn^{2+} and hydroxide ion, (b) Ag^+ and chromate ion, (c) Ba^{2+} and sulfite ion, (d) Rb^+ and sulfate ion, and (e) Li^+ and bicarbonate ion.

2.108 Write formulas for two compounds formed between O^{2-} and (a) lead, (b) tin, (c) manganese, (d) iron, and (e) copper.

2.109 Write formulas for the ionic compounds formed from Cl^- and (a) cadmium ion, (b) silver ion, (c) zinc ion, and (d) nickel ion.

Naming Compounds

2.110 Name the following molecular compounds.
(a) SiO_2 (b) XeF_4 (c) P_4O_{10} (d) Cl_2O_7

2.111 Name the following molecular compounds.
(a) ClF_3 (b) S_2Cl_2 (c) N_2O_5 (d) $AsCl_5$

2.112 Name the following ionic compounds.
(a) CaS (b) $AlBr_3$ (c) Na_3P (d) Ba_3As_2 (e) Rb_2S

2.113 Name the following ionic compounds.
(a) NaF (b) Mg_2C (c) Li_3N (d) Al_2O_3 (e) K_2Se

2.114 Name the following ionic compounds using the Stock system.
(a) FeS (b) CuO (c) SnO_2 (d) $CoCl_2 \cdot 6H_2O$

2.115 Name the following ionic compounds using the Stock system.
(a) Mn_2O_3 (b) Hg_2Cl_2 (c) PbS (d) $CrCl_3 \cdot 4H_2O$

2.116 Name the following. If necessary, refer to Table 2.8 on page 69.
(a) $NaNO_2$ (b) $KMnO_4$ (c) $MgSO_4 \cdot 7H_2O$ (d) KSCN

2.117 Name the following. If necessary, refer to Table 2.8 on page 69.
(a) K_3PO_4 (c) $Fe_2(CO_3)_3$
(b) $NH_4C_2H_3O_2$ (d) $Na_2S_2O_3 \cdot 5H_2O$

2.118 Identify each of the following as molecular or ionic and give its name:
(a) $CrCl_2$ (f) P_4O_6
(b) S_2Cl_2 (g) $CaSO_3$
(c) $NH_4C_2H_3O_2$ (h) AgCN
(d) SO_3 (i) $ZnBr_2$
(e) KIO_3 (j) H_2Se

2.119 Identify each of the following as molecular or ionic and give its name:
(a) $V(NO_3)_3$ (f) K_2CrO_4
(b) $Co(C_2H_3O_2)_2$ (g) $Fe(OH)_2$
(c) Au_2S_3 (h) I_2O_4
(d) Au_2S (i) I_4O_9
(e) $GeBr_4$ (j) P_4Se_3

OH **2.120** Write formulas for the following.
(a) sodium monohydrogen phosphate
(b) lithium selenide
(c) chromium(III) acetate
(d) disulfur decafluoride
(e) nickel(II) cyanide
(f) iron(III) oxide
(g) antimony pentafluoride

2.121 Write formulas for the following.
(a) dialuminum hexachloride
(b) tetraarsenic decaoxide
(c) magnesium hydroxide
(d) copper(II) bisulfate
(e) ammonium thiocyanate
(f) potassium thiosulfate
(g) diiodine pentaoxide

2.122 Write formulas for the following.
(a) ammonium sulfide
(b) chromium(III) sulfate hexahydrate
(c) silicon tetrafluoride
(d) molybdenum(IV) sulfide
(e) tin(IV) chloride
(f) hydrogen selenide
(g) tetraphosphorus heptasulfide

2.123 Write formulas for the following.
(a) mercury(II) acetate
(b) barium hydrogen sulfite
(c) boron trichloride
(d) calcium phosphide
(e) magnesium dihydrogen phosphate
(f) calcium oxalate
(g) xenon tetrafluoride

2.124 The compounds Se_2S_6 and Se_2S_4 have been shown to be antidandruff agents. What are their names?

2.125 The compound P_2S_5 is used to manufacture safety matches. What is the name of this compound?

ADDITIONAL EXERCISES

2.126 An element has 25 protons in its nucleus.
(a) Is the element a metal, a nonmetal, or a metalloid?
(b) On the basis of the average atomic mass, write the symbol for the element's most abundant isotope.
(c) How many neutrons are in the isotope you described in part (b)?
(d) How many electrons are in atoms of this element?
(e) How many times heavier than ^{12}C is the average atom of this element?

*__2.127__ Elements *X* and *Y* form a compound in which there is one atom of *X* for every four atoms of *Y*. When these elements react, it is found that 1.00 g of *X* combines with 5.07 g of *Y*. When 1.00 g of *X* combines with 1.14 g of O, it forms a compound containing two atoms of O for each atom of *X*. Calculate the atomic mass of *Y*.

2.128 An iron nail is composed of four isotopes with the percentage abundances and atomic masses given in the following table. Calculate the average atomic mass of iron.

Isotope	Percentage Abundance	Atomic Mass (u)
^{54}Fe	5.80	53.9396
^{56}Fe	91.72	55.9349
^{57}Fe	2.20	56.9354
^{58}Fe	0.28	57.9333

OH *__2.129__ Bromine (shown in Figure 2.12, page 53) is a dark red liquid that vaporizes easily and is very corrosive to the skin. It is used commercially as a bleach for fibers and silk. Naturally occurring bromine is composed of two isotopes: ^{79}Br, with a mass of 78.9183 u, and ^{81}Br, with a mass of 80.9163 u. Use this information and the average atomic mass of bromine given in the table on the inside front cover of the book to calculate the percentage abundances of the two isotopes.

OH *__2.130__ Rust contains an iron–oxygen compound in which there are three oxygen atoms for each two iron atoms. In this compound, the iron to oxygen mass ratio is 2.325 g Fe to 1.000 g O. Another compound of iron and oxygen contains these elements in the ratio of 2.616 g Fe to 1.000 g O. What is the ratio of iron to oxygen atoms in this other iron–oxygen compound?

2.131 One atomic mass unit has a mass of $1.6605389 \times 10^{-24}$ g. Calculate the mass, in grams, of one atom of magnesium. What is the mass of one atom of iron, expressed in grams? Use these two answers to determine how many atoms of Mg are in 24.31 g of magnesium and how many atoms of Fe are in 55.85 g of iron. Compare your answers. What conclusions can you draw from the results of these calculations? Without actually performing any calculations, how many atoms do you think would be in 40.08 g of calcium?

OH **2.132** What are the formulas for mercury(I) nitrate dihydrate and mercury(II) nitrate monohydrate?

2.133 Consider the following substances: Cl_2, CaO, HBr, $CuCl_2$, AsH_3, $NaNO_3$, and NO_2.
(a) Which are binary substances?
(b) Which is a triatomic molecule?
(c) In which do we find only electron sharing?
(d) Which are diatomic?
(e) In which do we find only attractions between ions?
(f) Which are molecular?
(g) Which are ionic?

2.134 Using the old system of nomenclature, write the names for the following compounds. (See Facets of Chemistry 2.3.)
(a) gold(III) sulfate
(b) gold(III) nitrate
(c) lead(IV) oxide
(d) mercury(I) chloride
(e) copper(II) sulfate
(f) mercury(II) chloride
(g) cobalt(II) hydroxide
(h) tin(II) chloride
(i) tin(IV) sulfide

2.135 Write the names of the following compounds using the Stock system of nomenclature. Also write their formulas. (See Facets of Chemistry 2.3.)
(a) cupric bromide
(b) cuprous iodide
(c) ferrous sulfate
(d) chromous chloride
(e) mercurous nitrate
(f) manganous sulfate
(g) plumbous acetate

2.136 A student needed a sample of $Fe(NO_3)_3 \cdot 9H_2O$, but when she went to the latest edition of the catalog of a major chemical supplier, she could not find it listed alphabetically under "iron." Knowing that suppliers of laboratory chemicals still often list chemicals under their older names, what name should she look for in the catalog?

2.137 Write the balanced chemical equation for the reaction between elements with atomic numbers of (a) 20 and 35, (b) 6 and 17, (c) 13 and 16. For each of these determine the ratio of the mass of the heavier element to the lighter element in the compound.

OH **2.138** Write the balanced gas phase chemical equation for the reaction of dinitrogen pentoxide with sulfur dioxide to form sulfur trioxide and nitrogen oxide. What small, whole-number ratios are expected for oxygen in the nitrogen oxides and the sulfur oxides?

EXERCISES IN CRITICAL THINKING

2.139 Imagine a world where, for some reason, hydrogen and helium have not been discovered. Would Mendeleev have had enough information to predict their existence?

2.140 Around 1750 Benjamin Franklin knew of two opposite types of electric charge, produced by rubbing a glass rod or amber rod with fur. He decided that the charge developed on the glass rod should be the "positive" charge and from there on charges were defined. What would have changed if Franklin decided the amber rod acquired the positive charge?

2.141 Explore the Internet and find for yourself a reliable source of physical properties of elements and compounds. Justify how you decided the site was reliable.

2.142 Spreadsheet applications such as Microsoft Excel can display data in a variety of ways; some of these are shown throughout this book. What method of displaying periodic trends (line graphs, tables, bar graphs, 3-D views, etc.) is most effective for your learning style? Explain your answer by stating why your chosen display is better than the others.

2.143 Scientists often validate measurements such as measuring the circumference of the earth by using two independent methods to measure the same value. Describe two independent methods for determining the atomic mass of an element. Explain how these methods are truly independent.