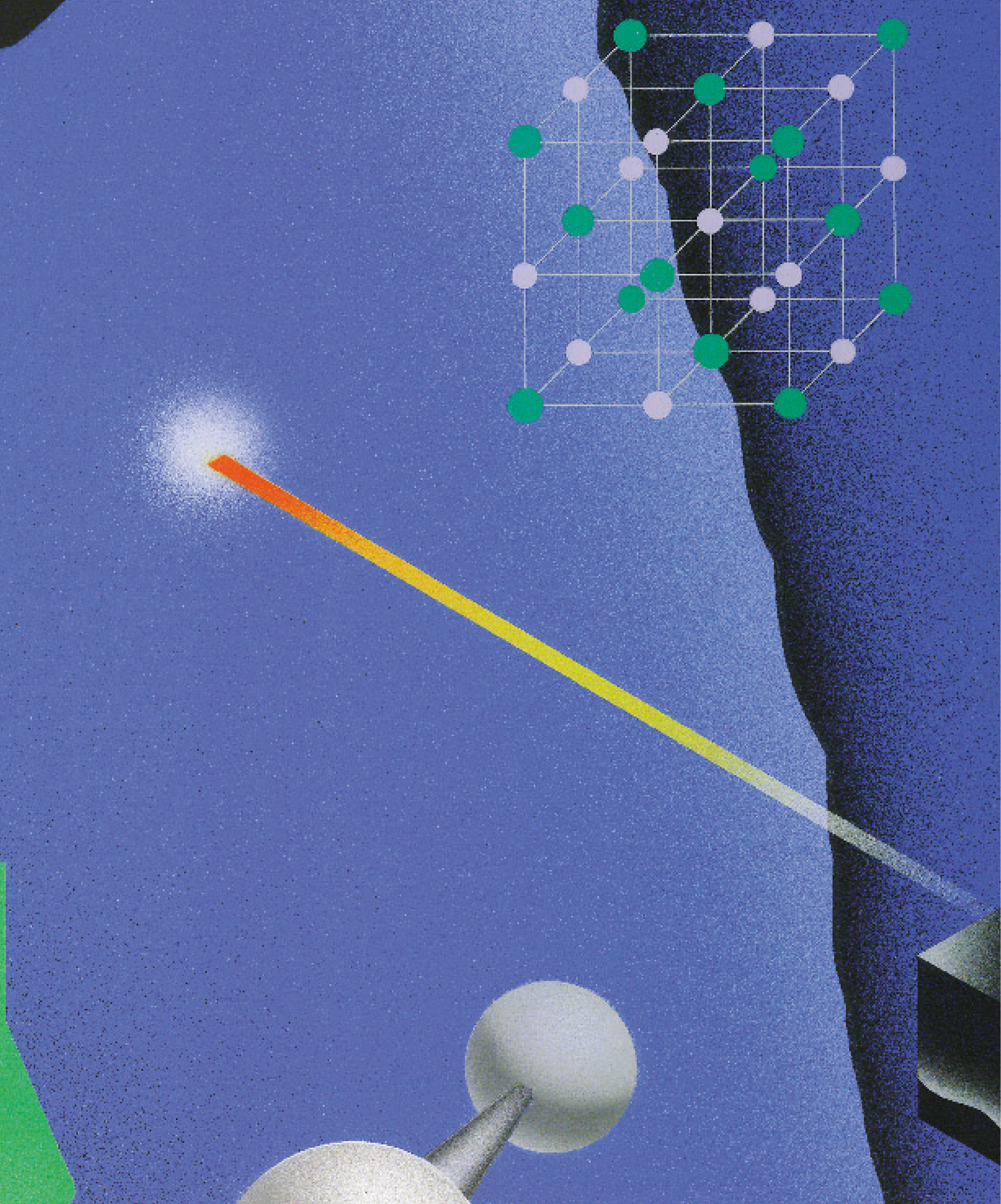


CHAPTER 2

Atoms, Molecules, and Ions

INTRODUCTION

SINCE ANCIENT TIMES HUMANS HAVE PONDERED THE NATURE OF MATTER. OUR MODERN IDEAS OF THE STRUCTURE OF MATTER BEGAN TO TAKE SHAPE IN THE EARLY NINETEENTH CENTURY WITH DALTON'S ATOMIC THEORY. WE NOW KNOW THAT ALL MATTER IS MADE OF ATOMS, MOLECULES, AND IONS. ALL OF CHEMISTRY IS CONCERNED IN ONE WAY OR ANOTHER WITH THESE SPECIES.

2.1 THE ATOMIC THEORY

In the fifth century B.C. the Greek philosopher Democritus expressed the belief that all matter consists of very small, indivisible particles, which he named *atomos* (meaning uncuttable or indivisible). Although Democritus' idea was not accepted by many of his contemporaries (notably Plato and Aristotle), somehow it endured. Experimental evidence from early scientific investigations provided support for the notion of "atomism" and gradually gave rise to the modern definitions of elements and compounds. It was in 1808 that an English scientist and school teacher, John Dalton,† formulated a precise definition of the indivisible building blocks of matter that we call atoms.

Dalton's work marked the beginning of the modern era of chemistry. The hypotheses about the nature of matter on which Dalton's atomic theory is based can be summarized as follows:

1. Elements are composed of extremely small particles called atoms. All atoms of a given element are identical, having the same size, mass, and chemical properties. The atoms of one element are different from the atoms of all other elements.
2. Compounds are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.
3. A chemical reaction involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

Figure 2.1 is a schematic representation of the first two hypotheses.

Dalton's concept of an atom was far more detailed and specific than Democritus'. The first hypothesis states that atoms of one element are different from atoms of all other elements. Dalton made no attempt to describe the structure or composition of atoms—he had no idea what an atom is really like. But he did realize that the different properties shown by elements such as hydrogen and oxygen can be explained by assuming that hydrogen atoms are not the same as oxygen atoms.

The second hypothesis suggests that, in order to form a certain compound, we need not only atoms of the right kinds of elements, but specific numbers of these atoms as well. This idea is an extension of a law published in 1799 by Joseph Proust,‡ a French chemist. Proust's ***law of definite proportions*** states that *different samples of the same compound always contain its constituent elements in the same proportion by mass.* Thus, if we were to analyze samples of carbon dioxide gas obtained from different sources, we would find in each sample the same ratio by mass of carbon to oxygen. It stands to reason, then, that if the ratio of the masses of different elements in a given compound is fixed, the ratio of the atoms of these elements in the compound also must be constant.

Dalton's second hypothesis supports another important law, the ***law of multiple proportions.*** According to the law, *if two elements can combine to form more than one*

†John Dalton (1766–1844). English chemist, mathematician, and philosopher. In addition to the atomic theory, he also formulated several gas laws and gave the first detailed description of color blindness, from which he suffered. Dalton was described as an indifferent experimenter, and singularly wanting in the language and power of illustration. His only recreation was lawn bowling on Thursday afternoons. Perhaps it was the sight of those wooden balls that provided him with the idea of the atomic theory.

‡Joseph Louis Proust (1754–1826). French chemist. Proust was the first person to isolate sugar from grapes.

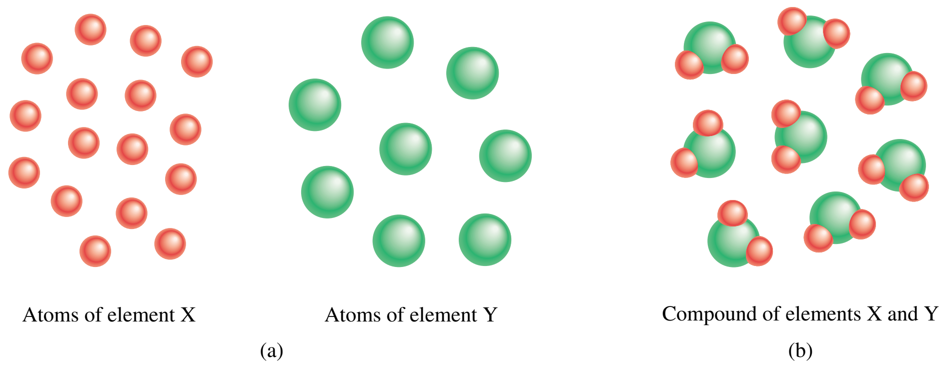


FIGURE 2.1 *(a) According to Dalton's atomic theory, atoms of the same element are identical, but atoms of one element are different from atoms of other elements. (b) Compound formed from atoms of elements X and Y. In this case, the ratio of the atoms of element X to the atoms of element Y is 2:1.*

compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers. Dalton's theory explains the law of multiple proportions quite simply: Different compounds made up of the same elements differ in the number of atoms of each kind that combine. For example, carbon forms two stable compounds with oxygen, namely, carbon monoxide and carbon dioxide. Modern measurement techniques indicate that one atom of carbon combines with one atom of oxygen in carbon monoxide and with two atoms of oxygen in carbon dioxide. Thus, the ratio of oxygen in carbon monoxide to oxygen in carbon dioxide is 1:2. This result is consistent with the law of multiple proportions.

Dalton's third hypothesis is another way of stating the ***law of conservation of mass,*** which is that *matter can be neither created nor destroyed.* Since matter is made of atoms that are unchanged in a chemical reaction, it follows that mass must be conserved as well. Dalton's brilliant insight into the nature of matter was the main stimulus for the rapid progress of chemistry during the nineteenth century.

2.2 THE STRUCTURE OF THE ATOM

On the basis of Dalton's atomic theory, we can define an ***atom*** as *the basic unit of an element that can enter into chemical combination.* Dalton imagined an atom that was both extremely small and indivisible. However, a series of investigations that began in the 1850s and extended into the twentieth century clearly demonstrated that atoms actually possess internal structure; that is, they are made up of even smaller particles, which are called *subatomic particles.* This research led to the discovery of three such particles—electrons, protons, and neutrons.

THE ELECTRON

In the 1890s many scientists became caught up in the study of ***radiation,*** *the emission and transmission of energy through space in the form of waves.* Information gained from this research contributed greatly to our understanding of atomic structure. One device used to investigate this phenomenon was a cathode ray tube, the forerunner of the television tube (Figure 2.2). It is a glass tube from which most of the air has been evacuated. When the two metal plates are connected to a high-voltage source, the negatively charged plate, called the *cathode,* emits an invisible ray. The cathode ray is

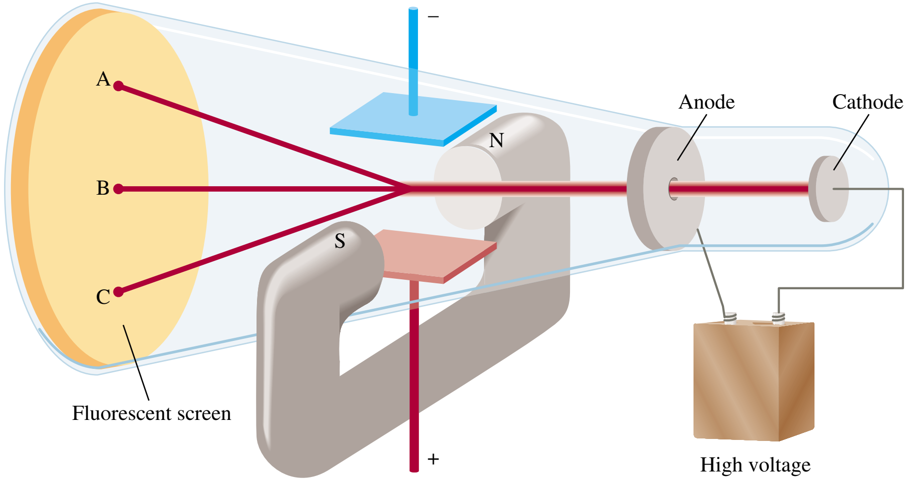


FIGURE 2.2 *A cathode ray tube with an electric field perpendicular to the direction of the cathode rays and an external magnetic field. The symbols N and S denote the north and south poles of the magnet. The cathode rays will strike the end of the tube at A in the presence of a magnetic field, at C in the presence of an electric field, and at B when there are no external fields present or when the effects of the electric field and magnetic field cancel each other.*

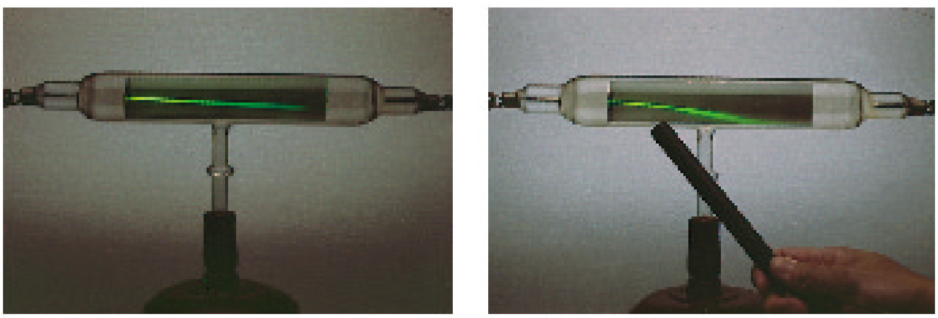
drawn to the positively charged plate, called the *anode,* where it passes through a hole and continues traveling to the other end of the tube. When the ray strikes the specially coated surface, it produces a strong fluorescence, or bright light.

In some experiments two electrically charged plates and a magnet were added to the *outside* of the cathode ray tube (see Figure 2.2). When the magnetic field is on and the electric field is off, the cathode ray strikes point A. When only the electric field is on, the ray strikes point C. When both the magnetic and the electric fields are off or when they are both on but balanced so that they cancel each other's influence, the ray strikes point B. According to electromagnetic theory, a moving charged body behaves like a magnet and can interact with electric and magnetic fields through which it passes. Since the cathode ray is attracted by the plate bearing positive charges and repelled by the plate bearing negative charges, it must consist of negatively charged particles. We know these *negatively charged particles* as ***electrons.*** Figure 2.3 shows the effect of a bar magnet on the cathode ray.

Electrons are normally associated with atoms. However, they can also be studied individually.

An English physicist, J. J. Thomson,[†] used a cathode ray tube and his knowledge of electromagnetic theory to determine the ratio of electric charge to the mass of an

[†]Joseph John Thomson (1856–1940). British physicist who received the Nobel Prize in Physics in 1906 for discovering the electron.



(a) (b)

FIGURE 2.3 *(a) A cathode ray produced in a discharge tube. The ray itself is invisible, but the fluorescence of a zinc sulfide coating on the glass causes it to appear green. (b) The cathode ray is bent in the presence of a magnet.*

individual electron. The number he came up with is -1.76×10^8 C/g, where C stands for *coulomb,* which is the unit of electric charge. Thereafter, in a series of experiments carried out between 1908 and 1917, R. A. Millikan[†] found the charge of an electron to be -1.60×10^{-19} C. From these data he calculated the mass of an electron:

$$\begin{aligned}\text{mass of an electron} &= \frac{\text{charge}}{\text{charge/mass}}\\ &= \frac{-1.60 \times 10^{-19}\ \text{C}}{-1.76 \times 10^{8}\ \text{C/g}}\\ &= 9.09 \times 10^{-28}\ \text{g}\end{aligned}$$

This is an exceedingly small mass.

RADIOACTIVITY

In 1895, the German physicist Wilhelm Röntgen[‡] noticed that cathode rays caused glass and metals to emit very unusual rays. This highly energetic radiation penetrated matter, darkened covered photographic plates, and caused a variety of substances to fluoresce. Since these rays could not be deflected by a magnet, they could not contain charged particles as cathode rays do. Röntgen called them X rays because their nature was not known.

Not long after Röntgen's discovery, Antoine Becquerel,[§] a professor of physics in Paris, began to study the fluorescent properties of substances. Purely by accident, he found that exposing thickly wrapped photographic plates to a certain uranium compound caused them to darken, even without the stimulation of cathode rays. Like X rays, the rays from the uranium compound were highly energetic and could not be deflected by a magnet, but they differed from X rays because they arose spontaneously. One of Becquerel's students, Marie Curie,[¶] suggested the name ***radioactivity*** to describe this *spontaneous emission of particles and/or radiation.* Since then, any element that spontaneously emits radiation is said to be *radioactive.*

Three types of rays are produced by the *decay,* or breakdown, of radioactive substances such as uranium. Two of the three are deflected by oppositely charged metal plates (Figure 2.4). ***Alpha (α) rays*** consist of *positively charged particles,* called ***α particles,*** and therefore are deflected by the positively charged plate. ***Beta (β) rays,*** or ***β particles,*** are *electrons* and are deflected by the negatively charged plate. The third type of radioactive radiation consists of high-energy rays called ***γ rays.*** Like X rays, γ rays have no charge and are not affected by an external field.

[†]Robert Andrews Millikan (1868–1953). American physicist who was awarded the Nobel Prize in Physics in 1923 for determining the charge of the electron.

[‡]Wilhelm Konrad Röntgen (1845–1923). German physicist who received the Nobel Prize in Physics in 1901 for the discovery of X rays.

[§]Antoine Henri Becquerel (1852–1908). French physicist who was awarded the Nobel Prize in Physics in 1903 for discovering radioactivity in uranium.

[¶]Marie (Marya Sklodowska) Curie (1867–1934). Polish-born chemist and physicist. In 1903 she and her French husband, Pierre Curie, were awarded the Nobel Prize in Physics for their work on radioactivity. In 1911, she again received the Nobel prize, this time in chemistry, for her work on the radioactive elements radium and polonium. She is one of only three people to have received two Nobel prizes in science. Despite her great contribution to science, her nomination to the French Academy of Sciences in 1911 was rejected by one vote because she was a woman! Her daughter Irene, and son-in-law Frederic Joliot-Curie, shared the Nobel Prize in Chemistry in 1935.

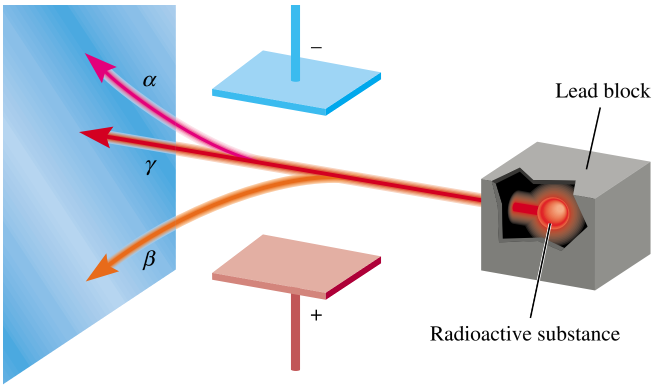


FIGURE 2.4 *Three types of rays emitted by radioactive elements. β Rays consist of negatively charged particles (electrons) and are therefore attracted by the positively charged plate. The opposite holds true for α rays—they are positively charged and are drawn to the negatively charged plate. Because γ rays have no charges, their path is unaffected by an external electric field.*

THE PROTON AND THE NUCLEUS

By the early 1900s, two features of atoms had become clear: they contain electrons, and they are electrically neutral. To maintain electric neutrality, an atom must contain an equal number of positive and negative charges. Therefore, Thomson proposed that an atom could be thought of as a uniform, positive sphere of matter in which electrons are embedded like raisins in a cake (Figure 2.5). This so-called "plum-pudding model" was the accepted theory for a number of years.

In 1910 the New Zealand physicist Ernest Rutherford,[†] who had studied with Thomson at Cambridge University, decided to use α particles to probe the structure of atoms. Together with his associate Hans Geiger[‡] and an undergraduate named Ernest Marsden,[§] Rutherford carried out a series of experiments using very thin foils of gold and other metals as targets for α particles from a radioactive source (Figure 2.6). They observed that the majority of particles penetrated the foil either undeflected or with only a slight deflection. But every now and then an α particle was scattered (or deflected) at a large angle. In some instances, an α particle actually bounced back in the direction from which it had come! This was a most surprising finding, for in Thomson's model the positive charge of the atom was so diffuse that the positive α particles should have passed through the foil with very little deflection. To quote Rutherford's initial reaction when told of this discovery: "It was as incredible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you."

Rutherford was later able to explain the results of the α-scattering experiment in terms of a new model for the atom. According to Rutherford, most of the atom must

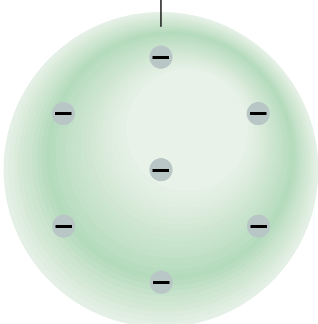


FIGURE 2.5 *Thomson's model of the atom, sometimes described as the "plum pudding" model, after a traditional English dessert containing raisins. The electrons are embedded in a uniform, positively charged sphere.*

[†]Ernest Rutherford (1871–1937). New Zealand physicist. Rutherford did most of his work in England (Manchester and Cambridge universities). He received the Nobel Prize in Chemistry in 1908 for his investigations into the structure of the atomic nucleus. His often-quoted comment to his students was that "all science is either physics or stamp-collecting."

[‡]Johannes Hans Wilhelm Geiger (1882–1945). German physicist. Geiger's work focused on the structure of the atomic nucleus and on radioactivity. He invented a device for measuring radiation that is now commonly called the Geiger counter.

[§]Ernest Marsden (1889–1970). English physicist. It is gratifying to know that at times an undergraduate can assist in winning a Nobel Prize. Marsden went on to contribute significantly to the development of science in New Zealand.

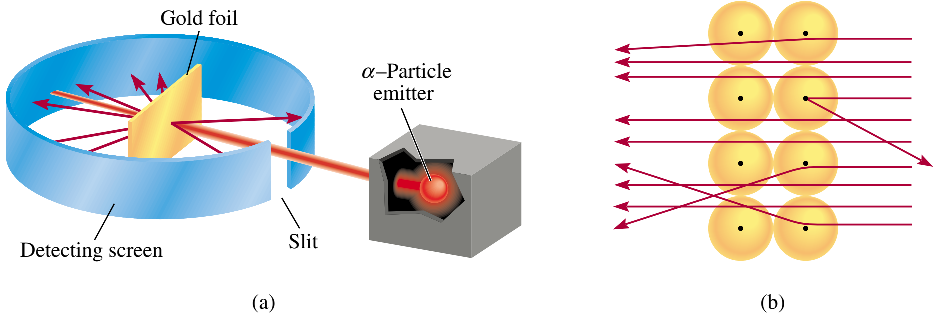


FIGURE 2.6 *(a) Rutherford's experimental design for measuring the scattering of α particles by a piece of gold foil. Most of the α particles passed through the gold foil with little or no deflection. A few were deflected at wide angles. Occasionally an α particle was turned back. (b) Magnified view of α particles passing through and being deflected by nuclei.*

be empty space. This explains why the majority of α particles passed through the gold foil with little or no deflection. The atom's positive charges, Rutherford proposed, are all concentrated in the ***nucleus,*** which is *a dense central core within the atom.* Whenever an α particle came close to a nucleus in the scattering experiment, it experienced a large repulsive force and therefore a large deflection. Moreover, an α particle traveling directly toward a nucleus would be completely repelled and its direction would be reversed.

The positively charged particles in the nucleus are called ***protons.*** In separate experiments, it was found that each proton carries the same *quantity* of charge as an electron and has a mass of 1.67252×10^{-24} g—about 1840 times the mass of the oppositely charged electron.

At this stage of investigation, scientists perceived the atom as follows: The mass of a nucleus constitutes most of the mass of the entire atom, but the nucleus occupies only about $1/10^{13}$ of the volume of the atom. We express atomic (and molecular) dimensions in terms of the SI unit called the *picometer* (*pm*), where

$$1 \text{ pm} = 1 \times 10^{-12} \text{ m}$$

A common non-SI unit for atomic length is the angstrom (Å; 1 Å = 100 pm).

A typical atomic radius is about 100 pm, whereas the radius of an atomic nucleus is only about 5×10^{-3} pm. You can appreciate the relative sizes of an atom and its nucleus by imagining that if an atom were the size of the Houston Astrodome, the volume of its nucleus would be comparable to that of a small marble. While the protons are confined to the nucleus of the atom, the electrons are conceived of as being spread out about the nucleus at some distance from it.

The concept of atomic radius is useful experimentally, but it should not be inferred that atoms have well-defined boundaries or surfaces. We will learn later that the outer regions of atoms are relatively "fuzzy."

THE NEUTRON

Rutherford's model of atomic structure left one major problem unsolved. It was known that hydrogen, the simplest atom, contains only one proton and that the helium atom contains two protons. Therefore, the ratio of the mass of a helium atom to that of a hydrogen atom should be 2:1. (Because electrons are much lighter than protons, their contribution to atomic mass can be ignored.) In reality, however, the ratio is 4:1. Rutherford and others postulated that there must be another type of subatomic particle in the atomic nucleus; the proof was provided by another English physicist, James

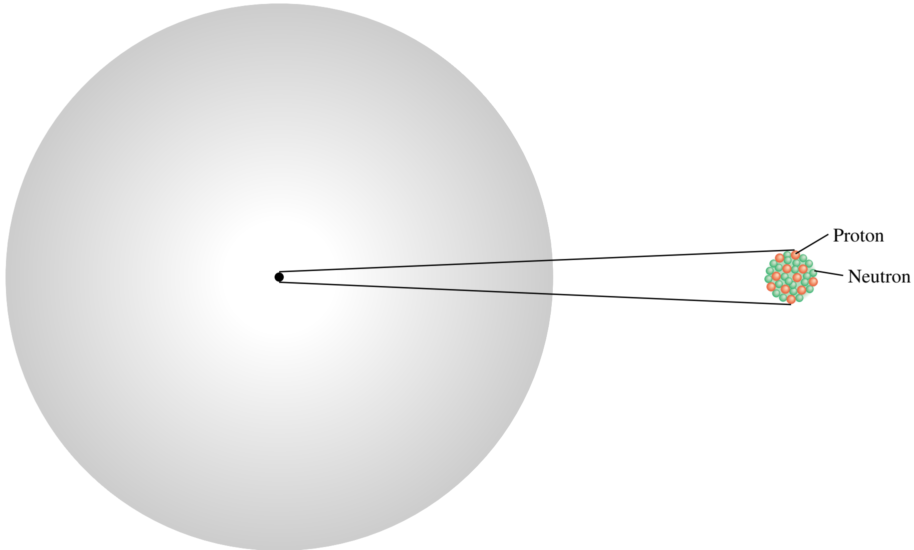


Figure 2.7 *The protons and neutrons of an atom are packed in an extremely small nucleus. Electrons are shown as "clouds" around the nucleus.*

Chadwick,† in 1932. When Chadwick bombarded a thin sheet of beryllium with α particles, a very high-energy radiation similar to γ rays was emitted by the metal. Later experiments showed that the rays actually consisted of a third type of subatomic particles, which Chadwick named ***neutrons,*** because they proved to be *electrically neutral particles having a mass slightly greater than that of protons.* The mystery of the mass ratio could now be explained. In the helium nucleus there are two protons and two neutrons, but in the hydrogen nucleus there is only one proton and no neutrons; therefore, the ratio is 4:1.

Figure 2.7 shows the location of the elementary particles (protons, neutrons, and electrons) in an atom. There are other subatomic particles, but the electron, the proton, and the neutron are the three fundamental components of the atom that are important in chemistry. Table 2.1 shows the masses and charges of these three elementary particles.

†James Chadwick (1891–1972). British physicist. In 1935 he received the Nobel Prize in Physics for proving the existence of neutrons.

TABLE 2.1 Mass and Charge of Subatomic Particles

		CHARGE	
PARTICLE	MASS (g)	COULOMB	CHARGE UNIT
Electron*	9.1095×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67252×10^{-24}	$+1.6022 \times 10^{-19}$	$+1$
Neutron	1.67495×10^{-24}	0	0

*More refined measurements have given us a more accurate value of an electron's mass than Millikan's.

2.3 ATOMIC NUMBER, MASS NUMBER, AND ISOTOPES

All atoms can be identified by the number of protons and neutrons they contain. The ***atomic number (Z)*** is *the number of protons in the nucleus of each atom of an element.* In a neutral atom the number of protons is equal to the number of electrons, so the atomic number also indicates the number of electrons present in the atom. The chemical identity of an atom can be determined solely from its atomic number. For example, the atomic number of nitrogen is 7. This means that each neutral nitrogen atom has 7 protons and 7 electrons. Or, viewed another way, every atom in the universe that contains 7 protons is correctly named "nitrogen."

The ***mass number (A)*** is *the total number of neutrons and protons present in the nucleus of an atom of an element.* Except for the most common form of hydrogen, which has one proton and no neutrons, all atomic nuclei contain both protons and neutrons. In general the mass number is given by

$$\begin{aligned}\text{mass number} &= \text{number of protons} + \text{number of neutrons}\\ &= \text{atomic number} + \text{number of neutrons}\end{aligned}$$

The number of neutrons in an atom is equal to the difference between the mass number and the atomic number, or $(A - Z)$. For example, the mass number of fluorine is 19 and the atomic number is 9 (indicating 9 protons in the nucleus). Thus the number of neutrons in an atom of fluorine is $19 - 9 = 10$. Note that the atomic number, number of neutrons, and mass number all must be positive integers (whole numbers).

Atoms of a given element do not all have the same mass. Most elements have two or more ***isotopes,*** *atoms that have the same atomic number but different mass numbers.* For example, there are three isotopes of hydrogen. One, simply known as hydrogen, has one proton and no neutrons. The *deuterium* isotope contains one proton and one neutron, and *tritium* has one proton and two neutrons. The accepted way to denote the atomic number and mass number of an atom of an element (X) is as follows:

mass number $\rightarrow$ $^{A}_{Z}\text{X}$

atomic number $\rightarrow$

Thus, for the isotopes of hydrogen, we write

$^{1}_{1}\text{H}$	$^{2}_{1}\text{H}$	$^{3}_{1}\text{H}$
hydrogen	deuterium	tritium

As another example, consider two common isotopes of uranium with mass numbers of 235 and 238, respectively:

$$^{235}_{92}\text{U} \qquad ^{238}_{92}\text{U}$$

The first isotope is used in nuclear reactors and atomic bombs, whereas the second isotope lacks the properties necessary for these applications. With the exception of hydrogen, which has different names for each of its isotopes, isotopes of elements are identified by their mass numbers. Thus the above two isotopes are called uranium-235 (pronounced "uranium two thirty-five") and uranium-238 (pronounced "uranium two thirty-eight").

The chemical properties of an element are determined primarily by the protons and electrons in its atoms; neutrons do not take part in chemical changes under normal conditions. Therefore, isotopes of the same element have similar chemistries, forming the same types of compounds and displaying similar reactivities.

The following example shows how to calculate the number of protons, neutrons, and electrons using atomic numbers and mass numbers.

EXAMPLE 2.1

Give the number of protons, neutrons, and electrons in each of the following species: (a) $^{17}_{8}O$, (b) $^{199}_{80}Hg$, (c) $^{200}_{80}Hg$.

Answer (a) The atomic number is 8, so there are 8 protons. The mass number is 17, so the number of neutrons is $17 - 8 = 9$. The number of electrons is the same as the number of protons, that is, 8.
(b) The atomic number is 80, so there are 80 protons. The mass number is 199, so the number of neutrons is $199 - 80 = 119$. The number of electrons is 80.
(c) Here the number of protons is the same as in (b), or 80. The number of neutrons is $200 - 80 = 120$. The number of electrons is also the same as in (b), 80. The species in (b) and (c) are chemically similar isotopes of mercury.

Similar problems: 2.15, 2.16.

PRACTICE EXERCISE
How many protons, neutrons, and electrons are in the following isotope of copper: $^{63}_{29}Cu$?

2.4 THE PERIODIC TABLE

More than half of the elements known today were discovered between 1800 and 1900. During this period, chemists noted that many elements show very strong similarities to one another. Recognition of periodic regularities in physical and chemical behavior and the need to organize the large volume of available information about the structure and properties of elemental substances led to the development of the ***periodic table,*** *a chart in which elements having similar chemical and physical properties are grouped together.* Figure 2.8 shows the modern periodic table in which the elements are arranged by atomic number (shown above the element symbol) in *horizontal rows* called ***periods*** and in *vertical columns* known as ***groups*** or ***families,*** according to similarities in their chemical properties. Note that Elements 110, 111, and 112 have recently been synthesized, although they have not yet been named.

The elements can be divided into three categories—metals, nonmetals, and metalloids. A ***metal*** is *a good conductor of heat and electricity* while a ***nonmetal*** is usually *a poor conductor of heat and electricity.* A ***metalloid*** *has properties that are intermediate between those of metals and nonmetals.* Figure 2.8 shows that the majority of known elements are metals; only seventeen elements are nonmetals, and eight elements are metalloids. From left to right across any period, the physical and chemical properties of the elements change gradually from metallic to nonmetallic.

Elements are often referred to collectively by their periodic table group number (Group 1A, Group 2A, and so on). However, for convenience, some element groups have been given special names. *The Group 1A elements (Li, Na, K, Rb, Cs, and Fr)* are called ***alkali metals,*** and *the Group 2A elements (Be, Mg, Ca, Sr, Ba, and Ra)* are called ***alkaline earth metals.*** *Elements in Group 7A (F, Cl, Br, I, and At)* are known as

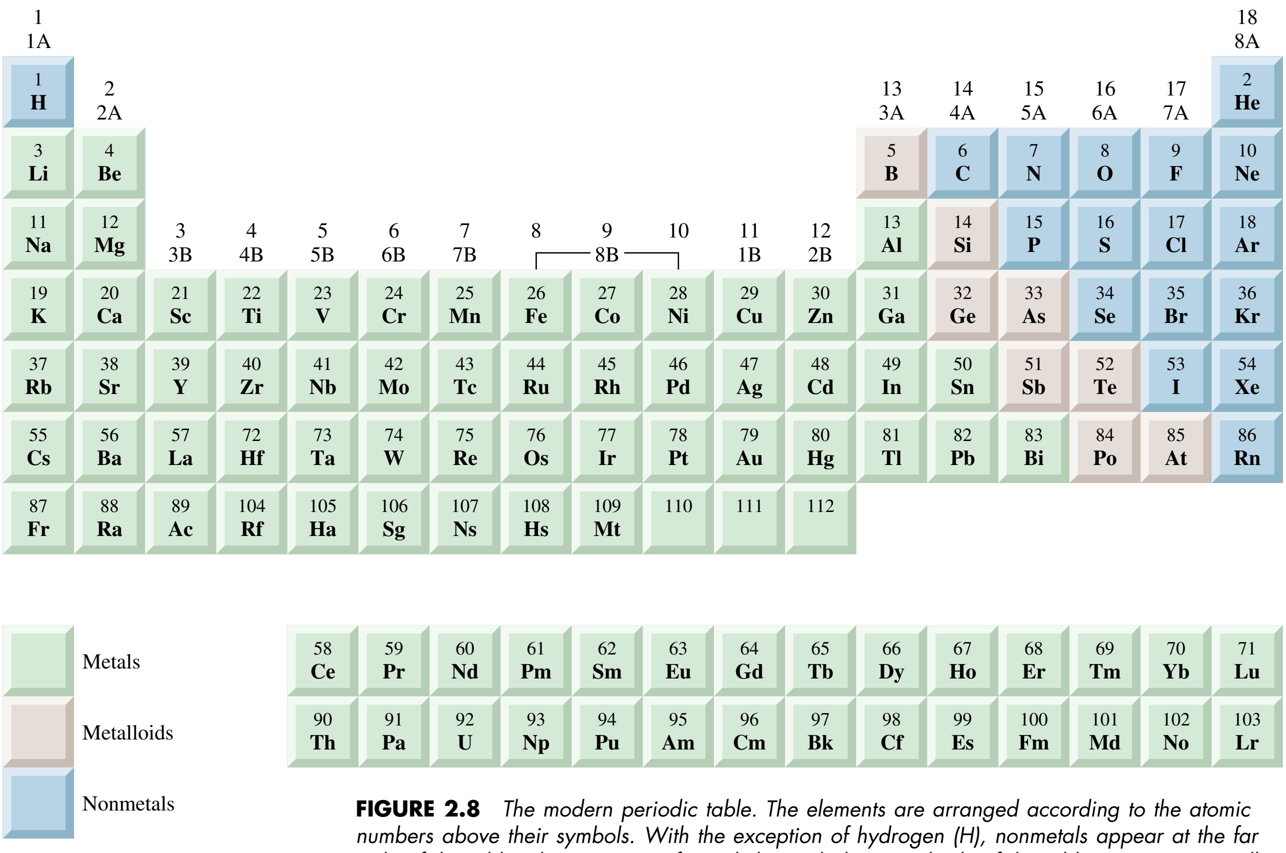


FIGURE 2.8 *The modern periodic table. The elements are arranged according to the atomic numbers above their symbols. With the exception of hydrogen (H), nonmetals appear at the far right of the table. The two rows of metals beneath the main body of the table are conventionally set apart to keep the table from being too wide. Actually, cerium (Ce) should follow lanthanum (La), and thorium (Th) should come right after actinium (Ac). The 1–18 group designation has been recommended by the International Union of Pure and Applied Chemistry (IUPAC) but is not yet in wide use. In this text we use the standard U.S. notation for group numbers (1A–8A and 1B–8B). The names for elements 104 to 109 are controversial. No names have yet been assigned for elements 110 to 112.*

halogens, and *elements in Group 8A (He, Ne, Ar, Kr, Xe, and Rn)* are called ***noble gases,*** or *rare gases.*

The periodic table is a handy tool that correlates the properties of the elements in a systematic way and helps us to make predictions about chemical behavior. We will take a closer look at this keystone of chemistry in Chapter 8.

The Chemistry in Action essay on p. 62 describes the distribution of the elements on Earth and in the human body.

2.5 MOLECULES AND IONS

Of all the elements, only the six noble gases in Group 8A of the periodic table (He, Ne, Ar, Kr, Xe, and Rn) exist in nature as single atoms. For this reason, they are called *monatomic* (meaning a single atom) gases. Most matter is composed of molecules or ions formed by atoms.

MOLECULES

We will discuss the nature of chemical bonds in Chapters 9 and 10.

A ***molecule*** is an *aggregate of at least two atoms in a definite arrangement held together by chemical forces* (also called *chemical bonds*). A molecule may contain atoms of the same element or atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions stated in Section 2.1. Thus, a molecule is not necessarily a compound, which, by definition, is made up of two or more elements (see Section 1.2). Hydrogen gas, for example, is a pure element, but it consists of molecules made up of two H atoms each. Water, on the other hand, is a molecular compound that contains hydrogen and oxygen in a ratio of two H atoms and one O atom. Like atoms, molecules are electrically neutral.

The hydrogen molecule, symbolized as H_2, is called a ***diatomic molecule*** because it *contains only two atoms.* Other elements that normally exist as diatomic molecules are nitrogen (N_2) and oxygen (O_2), as well as the Group 7A elements—fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). Of course, a diatomic molecule can contain atoms of different elements. Examples are hydrogen chloride (HCl) and carbon monoxide (CO).

The vast majority of molecules contain more than two atoms. They can be atoms of the same element, as in ozone (O_3), which is made up of three atoms of oxygen, or they can be combinations of two or more different elements. *Molecules containing more than two atoms* are called **polyatomic molecules.** Like ozone, water (H_2O) and ammonia (NH_3) are polyatomic molecules.

Molecular Models

Molecules are too small for us to observe directly. An effective means of visualizing them is by the use of molecular models. Two standard types of molecular models are currently in use: *ball-and-stick* models and *space-filling* models (Figure 2.9). In ball-and-stick model kits the atoms are wooden or plastic balls with holes in them. Sticks or springs are used to represent chemical bonds. The angles they form between atoms approximate the bond angles in actual molecules. The balls are all the same size and each type of atom is represented by a specific color. In space-filling models atoms are represented by truncated balls held together by snap fasteners, so that the bonds are not visible. The balls are proportional in size to atoms.

Ball-and-stick models show the three-dimensional arrangement of atoms clearly, and they are fairly easy to construct. However, the balls are not proportional to the size of atoms. Furthermore, the sticks greatly exaggerate the space between atoms in a molecule. Space-filling models are more accurate because they show the variation in atomic size. Their drawbacks are that they are time-consuming to put together, and they do not show the three-dimensional positions of atoms very well. We will use mostly the ball-and-stick model in this text.

IONS

An ***ion*** is *a charged species formed from a neutral atom or molecule when electrons are gained or lost as the result of a chemical change.* The number of positively charged protons in the nucleus of an atom remains the same during ordinary chemical changes (called chemical reactions), but negatively charged electrons may be lost or gained. The loss of one or more electrons from a neutral atom results in a ***cation,*** *an ion with*

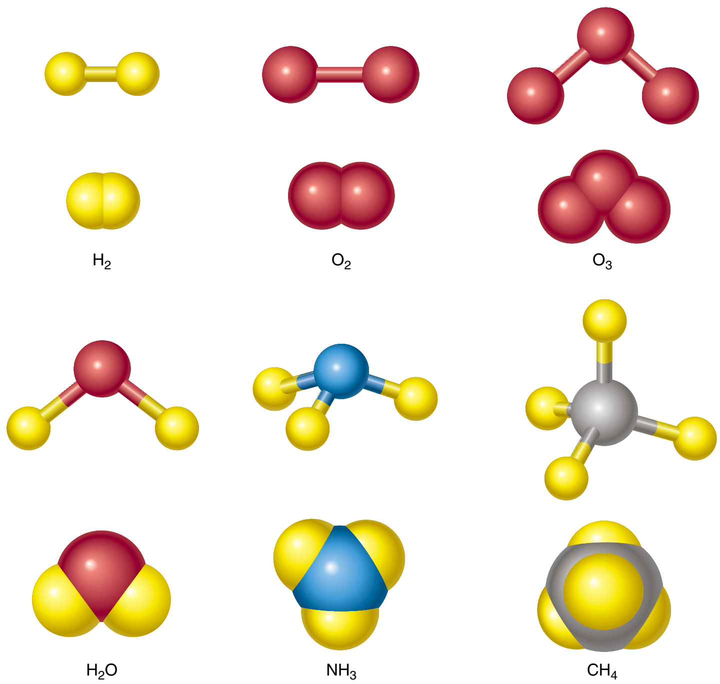


FIGURE 2.9 *Ball-and-stick and space-filling models of some simple molecules.*

In Chapter 8 we will see why atoms of different elements gain (or lose) a specific number of electrons.

a net positive charge. For example, a sodium atom (Na) can readily lose an electron to become sodium cation, which is represented by Na^+:

Na ATOM	Na^+ ION
11 protons	11 protons
11 electrons	10 electrons

On the other hand, an ***anion*** is *an ion whose net charge is negative* due to an increase in the number of electrons. A chlorine atom (Cl), for instance, can gain an electron to become the chloride ion Cl^-:

Cl ATOM	Cl^- ION
17 protons	17 protons
17 electrons	18 electrons

Sodium chloride (NaCl), ordinary table salt, is called an ***ionic compound*** because it is *formed from cations and anions.*

An atom can lose or gain more than one electron. Examples of ions formed by the loss or gain of more than one electron are Mg^{2+}, Fe^{3+}, S^{2-}, and N^{3-}. These ions, as well as Na^+ and Cl^-, are called ***monatomic ions*** because they *contain only one atom.* Figure 2.10 shows the charges of a number of monatomic ions. With very few exceptions, metals tend to form cations and nonmetals form anions.

In addition, two or more atoms can combine to form an ion that has a net positive or net negative charge. ***Polyatomic ions*** such as OH^- (hydroxide ion), CN^- (cyanide ion), and NH_4^+ (ammonium ion) are *ions containing more than one atom.*

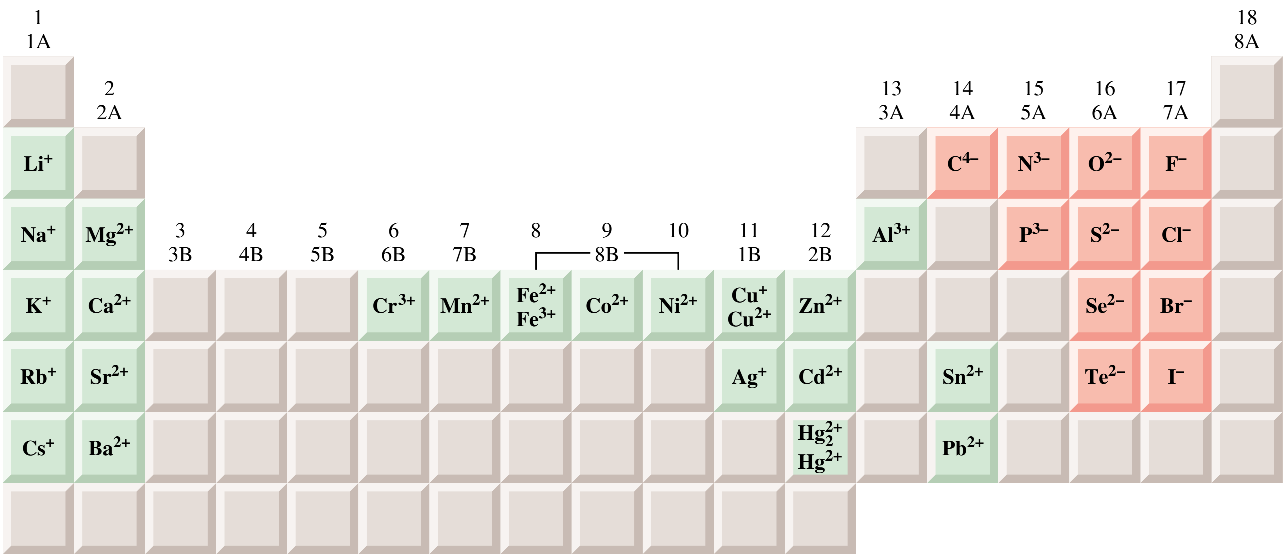


FIGURE 2.10 *Common monatomic ions arranged according to their positions in the periodic table. Note that the Hg_2^{2+} ion contains two atoms.*

2.6 CHEMICAL FORMULAS

Chemists use ***chemical formulas*** to *express the composition of molecules and ionic compounds in terms of chemical symbols.* By composition we mean not only the elements present but also the ratios in which the atoms are combined. Here we are concerned with two types of formulas: molecular formulas and empirical formulas.

MOLECULAR FORMULAS

A ***molecular formula*** *shows the exact number of atoms of each element in the smallest unit of a substance.* In our discussion of molecules, each example was given with its molecular formula in parentheses. Thus H_2 is the molecular formula for hydrogen, O_2 is oxygen, O_3 is ozone, and H_2O is water. The subscript numeral indicates the number of atoms of an element present. There is no subscript for O in H_2O because there is only one atom of oxygen in a molecule of water, and so the number "one" is omitted from the formula. Note that oxygen (O_2) and ozone (O_3) are allotropes of oxygen. An ***allotrope*** is *one of two or more distinct forms of an element.* Two allotropic forms of the element carbon—diamond and graphite—are dramatically different not only in properties but also in their relative cost.

EMPIRICAL FORMULAS

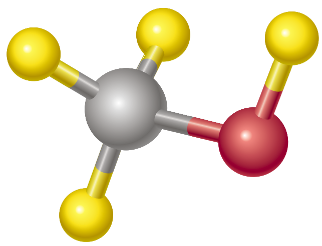
The molecular formula of hydrogen peroxide, a substance used as an antiseptic and as a bleaching agent for textiles and hair, is H_2O_2. This formula indicates that each hydrogen peroxide molecule consists of two hydrogen atoms and two oxygen atoms. The ratio of hydrogen to oxygen atoms in this molecule is 2:2 or 1:1. The empirical formula of hydrogen peroxide is HO. Thus the ***empirical formula*** *tells us which elements are present and the simplest whole-number ratio of their atoms,* but not necessarily the actual number of atoms in a given molecule. As another example, consider the compound hydrazine (N_2H_4), which is used as a rocket fuel. The empirical formula of hydrazine is NH_2. Although the ratio of nitrogen to hydrogen is 1:2 in both the molecular formula (N_2H_4) and the empirical formula (NH_2), only the molecular formula tells us the actual number of N atoms (two) and H atoms (four) present in a hydrazine molecule.

The word "empirical" means "derived from experiment." As we will see in Chapter 3, empirical formulas are determined experimentally.

Empirical formulas are the *simplest* chemical formulas; they are written by reducing the subscripts in molecular formulas to the smallest possible whole numbers. Molecular formulas are the *true* formulas of molecules. As we will see in Chapter 3, when chemists analyze an unknown compound, the first step is usually the determination of the compound's empirical formula.

For many molecules, the molecular formula and the empirical formula are one and the same. Some examples are water (H_2O), ammonia (NH_3), carbon dioxide (CO_2), and methane (CH_4).

The following two examples deal with writing molecular formulas from molecular models and writing empirical formulas from molecular formulas.



Methanol. (Gray=carbon, yellow=hydrogen, red=oxygen.)

EXAMPLE 2.2

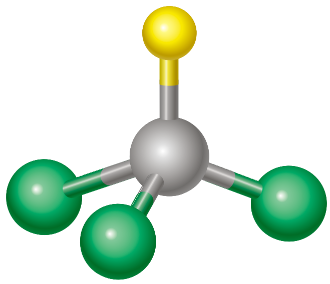
Write the molecular formula of methanol, an organic solvent and antifreeze, from its ball-and-stick model, shown in the margin.

Answer There is one C atom, four H atoms, and one O atom. Therefore, the molecular formula is CH_4O. However, the standard way of writing the molecular formula for methanol is CH_3OH because it shows how the atoms are joined in the molecule.

Similar problems: 2.41, 2.42.

PRACTICE EXERCISE

Write the molecular formula of chloroform, which is used as a solvent and a cleansing agent. The ball-and-stick model of chloroform is shown in the margin.



Chloroform. (Gray=carbon, yellow=hydrogen, green=chlorine.)

EXAMPLE 2.3

Write the empirical formulas for the following molecules: (a) acetylene (C_2H_2), which is used in welding torches; (b) glucose ($C_6H_{12}O_6$), a substance known as blood sugar; and (c) nitrous oxide (N_2O), a gas that is used as an anesthetic gas ("laughing gas") and as an aerosol propellant for whipped creams.

Answer (a) There are two carbon atoms and two hydrogen atoms in acetylene. Dividing the subscripts by 2, we obtain the empirical formula CH.

(b) In glucose there are six carbon atoms, twelve hydrogen atoms, and six oxygen atoms. Dividing the subscripts by 6, we obtain the empirical formula CH_2O. Note that if we had divided the subscripts by 3, we would have obtained the formula $C_2H_4O_2$. Although the ratio of carbon to hydrogen to oxygen atoms in $C_2H_4O_2$ is the same as that in $C_6H_{12}O_6$ (1:2:1), $C_2H_4O_2$ is not the simplest formula because its subscripts are not in the smallest whole-number ratio.

(c) Since the subscripts in N_2O are already the smallest possible whole numbers, the empirical formula for nitrous oxide is the same as its molecular formula.

Similar problems: 2.43, 2.44.

PRACTICE EXERCISE
Write the empirical formula for caffeine ($C_8H_{10}N_4O_2$), a stimulant found in tea and coffee.

The formulas of ionic compounds are always the same as their empirical formulas because ionic compounds do not consist of discrete molecular units. For example, a solid sample of sodium chloride (NaCl) consists of equal numbers of Na^+ and Cl^- ions arranged in a three-dimensional network (Figure 2.11). In such a compound there is a 1:1 ratio of cations to anions so that the compound is electrically neutral. As you can see in Figure 2.11, no Na^+ ion in NaCl is associated with just one particular Cl^- ion. In fact, each Na^+ ion is equally held by six surrounding Cl^- ions and vice versa. Thus NaCl is the empirical formula for sodium chloride. In other ionic compounds the actual structure may be different, but the arrangement of cations and anions is such that the compounds are all electrically neutral. Note that the charges on the cation and anion are not shown in the formula for an ionic compound.

In order for ionic compounds to be electrically neutral, the sum of the charges on the cation and anion in each formula unit must be zero. If the charges on the cation and anion are numerically different, we apply the following rule to make the formula electrically neutral: *The subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation.* If the charges are numerically equal, then no subscripts are necessary. This rule follows from the fact that because the formulas of ionic compounds are empirical formulas, the subscripts must always be reduced to the smallest ratios. Let us consider some examples.

- *Potassium bromide.* The potassium cation K^+ and the bromine anion Br^- combine to form the ionic compound potassium bromide. The sum of the charges is $+1 + (-1) = 0$, so no subscripts are necessary. The formula is KBr.
- *Zinc iodide.* The zinc cation Zn^{2+} and the iodine anion I^- combine to form zinc iodide. The sum of the charges of one Zn^{2+} ion and one I^- ion is $+2 + (-1) = +1$. To make the charges add up to zero we multiply the -1 charge of the anion by 2 and add the subscript "2" to the symbol for iodine. Therefore the formula for zinc iodide is ZnI_2.
- *Aluminum oxide.* The cation is Al^{3+} and the oxygen anion is O^{2-}. The following

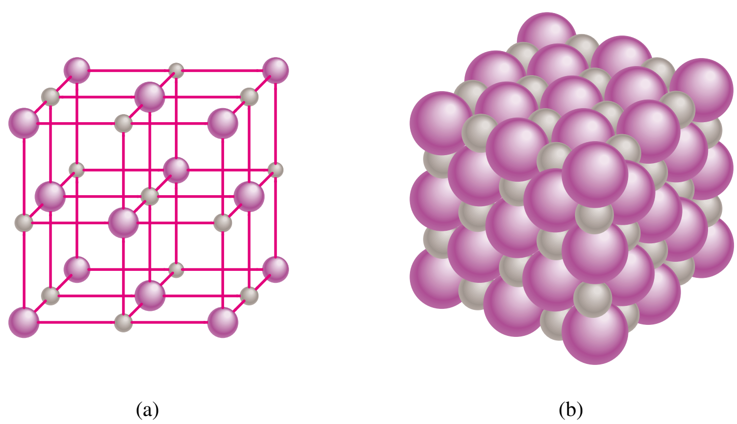


FIGURE 2.11 *(a) Structure of solid NaCl. (b) In reality, the cations are in contact with the anions. In both (a) and (b), the smaller spheres represent Na^+ ions and the larger spheres, Cl^- ions.*

diagram helps us determine the subscripts for the compound formed by the cation and the anion:

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

$$Al_2\, O_3$$

The sum of the charges is $2(+3) + 3(-2) = 0$. Thus the formula for aluminum oxide is Al_2O_3.

2.7 NAMING COMPOUNDS

When chemistry was a young science and the number of known compounds was small, it was possible to memorize their names. Many of the names were derived from their physical appearance, properties, origin, or application—for example, milk of magnesia, laughing gas, limestone, caustic soda, lye, washing soda, and baking soda.

Today the number of known compounds is well over 13 million. Fortunately, it is not necessary to memorize their names. Over the years chemists have devised a clear system for naming chemical substances. The rules are accepted worldwide, facilitating communication among chemists and providing a useful way of labeling an overwhelming variety of substances. Mastering these rules now will prove beneficial almost immediately as we proceed with our study of chemistry.

To begin our discussion of chemical *nomenclature,* the naming of chemical compounds, we must first distinguish between inorganic and organic compounds. *Organic compounds* contain carbon, usually in combination with elements such as hydrogen, oxygen, nitrogen, and sulfur. All other compounds are classified as *inorganic compounds.* For convenience, some carbon-containing compounds, such as carbon monoxide (CO), carbon dioxide (CO_2), carbon disulfide (CS_2), compounds containing the cyanide group (CN^-), and carbonate (CO_3^{2-}) and bicarbonate (HCO_3^-) groups are considered to be inorganic compounds. Although the nomenclature of organic compounds will not be discussed until Chapter 24, we will use some organic compounds to illustrate chemical principles throughout this book.

To organize and simplify our venture into naming compounds, we can divide inorganic compounds into four categories: ionic compounds, molecular compounds, acids and bases, and hydrates.

IONIC COMPOUNDS

In Section 2.5 we learned that ionic compounds are made up of cations (positive ions) and anions (negative ions). With the important exception of the ammonium ion, NH_4^+, all cations of interest to us are derived from metal atoms. Metal cations take their names from the elements. For example:

ELEMENT		NAME OF CATION	
Na	sodium	Na^+	sodium ion (or sodium cation)
K	potassium	K^+	potassium ion (or potassium cation)
Mg	magnesium	Mg^{2+}	magnesium ion (or magnesium cation)
Al	aluminum	Al^{3+}	aluminum ion (or aluminum cation)

Many ionic compounds are ***binary compounds,*** or *compounds formed from just two elements.* For binary compounds the first element named is the metal cation, followed by the nonmetallic anion. Thus NaCl is sodium chloride. The anion is named by taking the first part of the element name (chlorine) and adding "-ide." Potassium bromide (KBr), zinc iodide (ZnI_2), and aluminum oxide (Al_2O_3) are also binary compounds. Table 2.2 shows the "-ide" nomenclature of some common monatomic anions according to their positions in the periodic table.

The "-ide" ending is also used for certain anion groups containing different elements, such as hydroxide (OH^-) and cyanide (CN^-). Thus the compounds LiOH and KCN are named lithium hydroxide and potassium cyanide, respectively. These and a number of other such ionic substances are called ***ternary compounds,*** meaning *compounds consisting of three elements.* Table 2.3 lists alphabetically the names of a number of common cations and anions.

The transition metals are the elements in Groups 1B and 3B-8B (see Figure 2.8).

Certain metals, especially the *transition metals,* can form more than one type of cation. Take iron as an example. Iron can form two cations: Fe^{2+} and Fe^{3+}. An older nomenclature system that is still in limited use assigns the ending "-ous" to the cation with fewer positive charges and the ending "-ic" to the cation with more positive charges:

Fe^{2+}	ferrous ion
Fe^{3+}	ferric ion

The names of the compounds that these iron ions form with chlorine would thus be

$FeCl_2$	ferrous chloride
$FeCl_3$	ferric chloride

$FeCl_2$ (left) and $FeCl_3$ (right).

This method of naming ions has some distinct limitations. First, the "-ous" and "-ic" suffixes do not provide information regarding the actual charges of the two cations involved. Thus the ferric ion is Fe^{3+}, but the cation of copper named cupric has the formula Cu^{2+}. In addition, the "-ous" and "-ic" designations provide names for only two different elemental cations. Some metallic elements can assume three or more different positive charges in compounds. Therefore, it has become increasingly common to designate different cations with Roman numerals. This is called the Stock† system. In this system, the Roman numeral I indicates one positive charge, II means two positive charges, and so on. For example, manganese (Mn) atoms can assume several different positive charges:

Mn^{2+}: MnO	manganese(II) oxide
Mn^{3+}: Mn_2O_3	manganese(III) oxide
Mn^{4+}: MnO_2	manganese(IV) oxide

These names are pronounced "manganese-two oxide," "manganese-three oxide," and "manganese-four oxide." Using the Stock system, we denote the ferrous ion and the ferric ion as iron(II) and iron(III), respectively; ferrous chloride becomes iron(II) chloride; and ferric chloride is called iron(III) chloride. In keeping with modern practice, we will favor the Stock system of naming compounds in this textbook.

The following examples illustrate how to name ionic compounds and write formulas for ionic compounds based on the information given in Figure 2.10 and Tables 2.2 and 2.3.

†Alfred E. Stock (1876–1946). German chemist. Stock did most of his research in the synthesis and characterization of boron, beryllium, and silicon compounds. He was the first scientist to explore the dangers of mercury poisoning.

TABLE 2.2 The "-ide" Nomenclature of Some Common Monatomic Anions According to Their Positions in the Periodic Table

GROUP 4A	GROUP 5A	GROUP 6A	GROUP 7A
C Carbide (C^{4-})*	N Nitride (N^{3-})	O Oxide (O^{2-})	F Fluoride (F^-)
Si Silicide (Si^{4-})	P Phosphide (P^{3-})	S Sulfide (S^{2-})	Cl Chloride (Cl^-)
		Se Selenide (Se^{2-})	Br Bromide (Br^-)
		Te Telluride (Te^{2-})	I Iodide (I^-)

*The word "carbide" is also used for the anion C_2^{2-}.

TABLE 2.3 Names and Formulas of Some Common Inorganic Cations and Anions

CATION	ANION
Aluminum (Al^{3+})	Bromide (Br^-)
Ammonium (NH_4^+)	Carbonate (CO_3^{2-})
Barium (Ba^{2+})	Chlorate (ClO_3^-)
Cadmium (Cd^{2+})	Chloride (Cl^-)
Calcium (Ca^{2+})	Chromate (CrO_4^{2-})
Cesium (Cs^+)	Cyanide (CN^-)
Chromium(III) or chromic (Cr^{3+})	Dichromate ($Cr_2O_7^{2-}$)
Cobalt(II) or cobaltous (Co^{2+})	Dihydrogen phosphate ($H_2PO_4^-$)
Copper(I) or cuprous (Cu^+)	Fluoride (F^-)
Copper(II) or cupric (Cu^{2+})	Hydride (H^-)
Hydrogen (H^+)	Hydrogen carbonate or bicarbonate (HCO_3^-)
Iron(II) or ferrous (Fe^{2+})	Hydrogen phosphate (HPO_4^{2-})
Iron(III) or ferric (Fe^{3+})	Hydrogen sulfate or bisulfate (HSO_4^-)
Lead(II) or plumbous (Pb^{2+})	Hydroxide (OH^-)
Lithium (Li^+)	Iodide (I^-)
Magnesium (Mg^{2+})	Nitrate (NO_3^-)
Manganese(II) or manganous (Mn^{2+})	Nitride (N^{3-})
Mercury(I) or mercurous (Hg_2^{2+})*	Nitrite (NO_2^-)
Mercury(II) or mercuric (Hg^{2+})	Oxide (O^{2-})
Potassium (K^+)	Permanganate (MnO_4^-)
Silver (Ag^+)	Peroxide (O_2^{2-})
Sodium (Na^+)	Phosphate (PO_4^{3-})
Strontium (Sr^{2+})	Sulfate (SO_4^{2-})
Tin(II) or stannous (Sn^{2+})	Sulfide (S^{2-})
Zinc (Zn^{2+})	Sulfite (SO_3^{2-})
	Thiocyanate (SCN^-)

*Mercury(I) exists as a pair as shown.

EXAMPLE 2.4

Name the following ionic compounds: (a) $Cu(NO_3)_2$, (b) KH_2PO_4, and (c) NH_4ClO_3.

Answer (a) Since the nitrate ion (NO_3^-) bears one negative charge (see Table 2.3), the copper ion must have two positive charges. Therefore, the compound is copper(II) nitrate.

(b) The cation is K^+ and the anion is $H_2PO_4^-$ (dihydrogen phosphate). Since potassium only forms one type of ion (K^+), there is no need to use potassium(I) in the name. The compound is potassium dihydrogen phosphate.
(c) The cation is NH_4^+ (ammonium ion) and the anion is ClO_3^-. The compound is ammonium chlorate.

Similar problems: 2.53(a), (b), (e).

PRACTICE EXERCISE
Name the following compounds: (a) PbO, (b) Li_2SO_3.

EXAMPLE 2.5

Write chemical formulas for the following compounds: (a) mercury(I) nitrite, (b) cesium sulfide, and (c) calcium phosphate.

Answer: (a) The mercury(I) ion is diatomic, namely, Hg_2^{2+} (see Table 2.3), and the nitrite ion is NO_2^-. Therefore, the formula is $Hg_2(NO_2)_2$.
(b) Each sulfide ion bears two negative charges, and each cesium ion bears one positive charge (cesium is in Group 1A, as is sodium). Therefore, the formula is Cs_2S.
(c) Each calcium ion (Ca^{2+}) bears two positive charges, and each phosphate ion (PO_4^{3-}) bears three negative charges. To make the sum of the charges equal zero, we must adjust the numbers of cations and anions:

$$3(+2) + 2(-3) = 0$$

Thus the formula is $Ca_3(PO_4)_2$.

Similar problems: 2.55(a), (b), (h).

PRACTICE EXERCISE
Write formulas for the following ionic compounds: (a) rubidium sulfate, (b) barium hydride.

MOLECULAR COMPOUNDS

Unlike ionic compounds, molecular compounds contain discrete molecular units. They are usually composed of nonmetallic elements (see Figure 2.8). Many molecular compounds are binary compounds. Naming binary molecular compounds is similar to naming binary ionic compounds. We place the name of the first element in the formula first, and the second element is named by adding -ide to the root of the element name. Some examples are

HCl	hydrogen chloride
HBr	hydrogen bromide
SiC	silicon carbide

It is quite common for one pair of elements to form several different compounds. In these cases, confusion in naming the compounds is avoided by the use of Greek prefixes to denote the number of atoms of each element present (see Table 2.4). Consider the following examples:

CO	carbon monoxide
CO_2	carbon dioxide

TABLE 2.4 Greek Prefixes Used in Naming Molecular Compounds

PREFIX	MEANING
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

SO_2	sulfur dioxide
SO_3	sulfur trioxide
NO_2	nitrogen dioxide
N_2O_4	dinitrogen tetroxide

The following guidelines are helpful in naming compounds with prefixes:

- The prefix "mono-" may be omitted for the first element. For example, PCl_3 is named phosphorus trichloride, not monophosphorus trichloride. Thus the absence of a prefix for the first element usually means there is only one atom of that element present in the molecule.
- For oxides, the ending "a" in the prefix is sometimes omitted. For example, N_2O_4 may be called dinitrogen tetroxide rather than dinitrogen tetraoxide.

Exceptions to the use of Greek prefixes are molecular compounds containing hydrogen. Traditionally, many of these compounds are called either by their common, nonsystematic names or by names that do not specifically indicate the number of H atoms present:

B_2H_6	diborane
CH_4	methane
SiH_4	silane
NH_3	ammonia
PH_3	phosphine
H_2O	water
H_2S	hydrogen sulfide

Note that even the order of writing the elements in the formulas for hydrogen compounds is irregular. In water and hydrogen sulfide, H is written first, whereas it appears last in the other compounds.

Writing formulas for molecular compounds is usually straightforward. Thus the name arsenic trifluoride means that there are one As atom and three F atoms in each molecule, and the molecular formula is AsF_3. Note that the order of elements in the formula is the same as in its name.

EXAMPLE 2.6

Name the following molecular compounds: (a) $SiCl_4$ and (b) P_4O_{10}.

Answer (a) Since there are four chlorine atoms present, the compound is silicon tetrachloride.
(b) There are four phosphorus atoms and ten oxygen atoms present, so the compound is tetraphosphorus decoxide. Note that the "a" is omitted in "deca."

Similar problems: 2.53(c), (h), (j).

PRACTICE EXERCISE
Name the following molecular compounds: (a) NF_3 and (b) Cl_2O_7.

EXAMPLE 2.7

Write chemical formulas for the following molecular compounds: (a) carbon disulfide and (b) disilicon hexabromide.

Answer (a) Since there are one carbon atom and two sulfur atoms present, the formula is CS_2.
(b) There are two silicon atoms and six bromine atoms present, so the formula is Si_2Br_6.

Similar problems: 2.55(g), (j).

PRACTICE EXERCISE
Write chemical formulas for the following molecular compounds: (a) sulfur tetrafluoride, (b) dinitrogen pentoxide.

ACIDS AND BASES

Naming Acids

H^+ is equivalent to one *proton,* and is sometimes referred to that way.

An ***acid*** can be described as *a substance that yields hydrogen ions (H^+) when dissolved in water.* Formulas for acids contain one or more hydrogen atoms as well as an anionic group. Anions whose names end in "-ide" form acids with a "hydro-" prefix and an "-ic" ending, as shown in Table 2.5. In some cases two different names seem to be assigned to the same chemical formula.

HCl hydrogen chloride
HCl hydrochloric acid

The name assigned to the compound depends on its physical state. In the gaseous or pure liquid state, HCl is a molecular compound called hydrogen chloride. When it is dissolved in water, the molecules break up into H^+ and Cl^- ions; in this state, the substance is called hydrochloric acid.

Oxoacids are acids that *contain hydrogen, oxygen, and another element (the central element).* The formulas of oxoacids are usually written with the H first, followed by the central element and then O, as illustrated by the following examples:

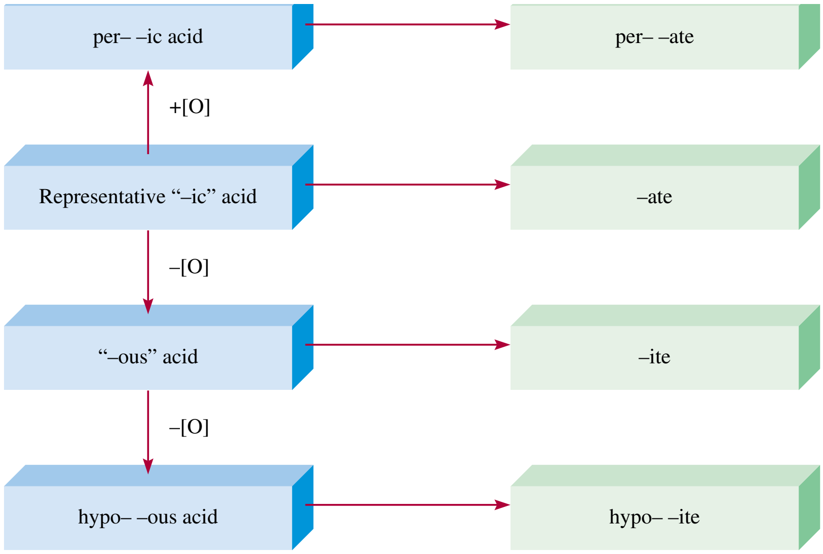
H_2CO_3 carbonic acid
HNO_3 nitric acid
H_2SO_4 sulfuric acid
$HClO_3$ chloric acid

Often two or more oxoacids have the same central atom but a different number of O atoms. Starting with the oxoacids whose names end with "-ic," we use the following rules to name these compounds.

TABLE 2.5 Some Simple Acids

ANION	CORRESPONDING ACID
F^- (fluoride)	HF (hydrofluoric acid)
Cl^- (chloride)	HCl (hydrochloric acid)
Br^- (bromide)	HBr (hydrobromic acid)
I^- (iodide)	HI (hydroiodic acid)
CN^- (cyanide)	HCN (hydrocyanic acid)
S^{2-} (sulfide)	H_2S (hydrosulfuric acid)

FIGURE 2.12 *Naming oxoacids and oxoanions.*



- Addition of one O atom to the "-ic" acid: The acid is called "per...-ic" acid. Thus adding an O atom to $HClO_3$ changes chloric acid to perchloric acid, $HClO_4$.
- Removal of one O atom from the "-ic" acid: The acid is called "-ous" acid. Thus nitric acid, HNO_3, becomes nitrous acid, HNO_2.
- Removal of two O atoms from the "-ic" acid: The acid is called "hypo...-ous" acid. Thus when $HBrO_3$ is converted to HBrO, the acid is called hypobromous acid.

The rules for naming ***oxoanions,*** anions of oxoacids, are as follows:

- When all the H ions are removed from the "-ic" acid, the anion's name ends with "-ate." For example, the anion CO_3^{2-} derived from H_2CO_3 is called carbonate.
- When all the H ions are removed from the "-ous" acid, the anion's name ends with "-ite." Thus the anion ClO_2^- derived from $HClO_2$ is called chlorite.
- The names of anions in which one or more but not all the hydrogen ions have been removed must indicate the number of H ions present. For example, consider the anions derived from phosphoric acid:

H_3PO_4	phosphoric acid
$H_2PO_4^-$	dihydrogen phosphate
HPO_4^{2-}	hydrogen phosphate
PO_4^{3-}	phosphate

Note that we usually omit the prefix "mono-" when there is only one H in the anion. Figure 2.12 summarizes the nomenclature for the oxoacids and oxoanions, and Table 2.6 gives the names of the oxoacids and oxoanions that contain chlorine.

TABLE 2.6 Names of Oxoacids and Oxoanions That Contain Chlorine

ACID	ANION
$HClO_4$ (perchloric acid)	ClO_4^- (perchlorate)
$HClO_3$ (chloric acid)	ClO_3^- (chlorate)
$HClO_2$ (chlorous acid)	ClO_2^- (chlorite)
$HClO$ (hypochlorous acid)	ClO^- (hypochlorite)

The following example deals with the nomenclature for an oxoacid and an oxanion.

EXAMPLE 2.8

Name the following oxoacid and oxoanion: (a) H_3PO_3, (b) IO_4^-.

Answer (a) We start with our reference acid, phosphoric acid (H_3PO_4). Since H_3PO_3 has one fewer O atom, it is called phosphorous acid.
(b) The parent acid is HIO_4. Since the acid has one more O atom than our reference iodic acid (HIO_3), it is called periodic acid. Therefore, the anion derived from HIO_4 is called periodate.

Similar problems: 2.53(h), 2.54(c).

PRACTICE EXERCISE
Name the following oxoacid and oxoanion: (a) HBrO, (b) HSO_4^-.

Naming Bases

A ***base*** can be described as *a substance that yields hydroxide ions (OH^-) when dissolved in water.* Some examples are

NaOH	sodium hydroxide
KOH	potassium hydroxide
$Ba(OH)_2$	barium hydroxide

Ammonia (NH_3), a molecular compound in the gaseous or pure liquid state, is also classified as a common base. At first glance this may seem to be an exception to the definition of a base. But note that as long as a substance *yields* hydroxide ions when dissolved in water, it need not contain hydroxide ions in its structure to be considered a base. In fact, when ammonia dissolves in water, NH_3 reacts partially with water to yield NH_4^+ and OH^- ions. Thus it is properly classified as a base.

HYDRATES

Hydrates are *compounds that have a specific number of water molecules attached to them.* For example, in its normal state, each unit of copper(II) sulfate has five water molecules associated with it. The systematic name for this compound is copper(II) sulfate pentahydrate, and its formula is written as $CuSO_4 \cdot 5H_2O$. The water molecules can be driven off by heating. When this occurs, the resulting compound is $CuSO_4$, which is sometimes called *anhydrous* copper(II) sulfate; "anhydrous" means that the compound no longer has water molecules associated with it (Figure 2.13). Some other

FIGURE 2.13 *$CuSO_4 \cdot 5H_2O$ (left) is blue; $CuSO_4$ (right) is white.*

TABLE 2.7 Common and Systematic Names of Some Compounds

FORMULA	COMMON NAME	SYSTEMATIC NAME
H_2O	Water	Dihydrogen oxide
NH_3	Ammonia	Trihydrogen nitride
CO_2	Dry ice	Solid carbon dioxide
NaCl	Table salt	Sodium chloride
N_2O	Laughing gas	Dinitrogen oxide (nitrous oxide)
$CaCO_3$	Marble, chalk, limestone	Calcium carbonate
CaO	Quicklime	Calcium oxide
$Ca(OH)_2$	Slaked lime	Calcium hydroxide
$NaHCO_3$	Baking soda	Sodium hydrogen carbonate
$Na_2CO_3 \cdot 10H_2O$	Washing soda	Sodium carbonate decahydrate
$MgSO_4 \cdot 7H_2O$	Epsom salt	Magnesium sulfate heptahydrate
$Mg(OH)_2$	Milk of magnesia	Magnesium hydroxide
$CaSO_4 \cdot 2H_2O$	Gypsum	Calcium sulfate dihydrate

hydrates are

$BaCl_2 \cdot 2H_2O$	barium chloride dihydrate
$LiCl \cdot H_2O$	lithium chloride monohydrate
$MgSO_4 \cdot 7H_2O$	magnesium sulfate heptahydrate
$Sr(NO_3)_2 \cdot 4H_2O$	strontium nitrate tetrahydrate

FAMILIAR INORGANIC COMPOUNDS

Some compounds are better known by their common names than by their systematic

SUMMARY OF FACTS AND CONCEPTS

1. Modern chemistry began with Dalton's atomic theory, which states that all matter is composed of tiny, indivisible particles called atoms; that all atoms of the same element are identical; that compounds contain atoms of different elements combined in whole-number ratios; and that atoms are neither created nor destroyed in chemical reactions (the law of conservation of mass).
2. Atoms of constituent elements in a particular compound are always combined in the same proportions by mass (law of definite proportions). When two elements can combine to form more than one type of compound, the masses of one element that combine with a fixed mass of the other element are in a ratio of small whole numbers (law of multiple proportions).

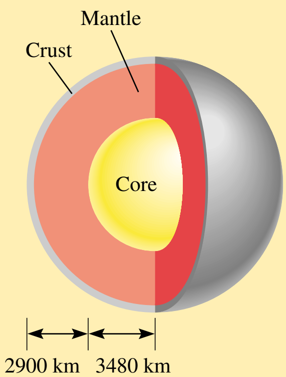
Distribution of Elements on Earth and in Living Systems

The majority of elements are naturally occurring. How are these elements distributed on Earth, and which are essential to living systems?

Earth's crust extends from the surface to a depth of about 40 km (about 25 miles). Because of technical difficulties, scientists have not been able to study the inner portions of Earth as easily as the crust. Nevertheless, it is believed that there is a solid core consisting mostly of iron at the center of Earth. Surrounding the core is a layer called the *mantle,* which consists of hot fluid containing iron, carbon, silicon, and sulfur.

Of the 83 elements that are found in nature, 12 of them make up 99.7 percent of Earth's crust by mass. They are, in decreasing order of natural abundance, oxygen (O), silicon (Si), aluminum (Al), iron (Fe), calcium (Ca), magnesium (Mg), sodium (Na), potassium (K), titanium (Ti), hydrogen (H), phosphorus (P), and manganese (Mn). In discussing the natural abundance of the elements, we should keep in mind that (1) the elements are not evenly distributed throughout Earth's crust, and (2) most elements occur in combined forms. These facts provide the basis for most methods of obtaining pure elements from their compounds, as we will see in later chapters.

The accompanying table lists the essential elements in the human body. Of special interest are the *trace elements,* such as iron (Fe), copper (Cu), zinc (Zn), iodine (I), and cobalt (Co), which together make up about one percent of the body's mass. These elements are necessary for biological functions such as growth, transport of oxygen for metabolism, and defense against disease. There is a delicate balance in the amounts of these elements in our bodies. Too much or too little over an extended period of time can lead to serious illness, retardation, or even death.



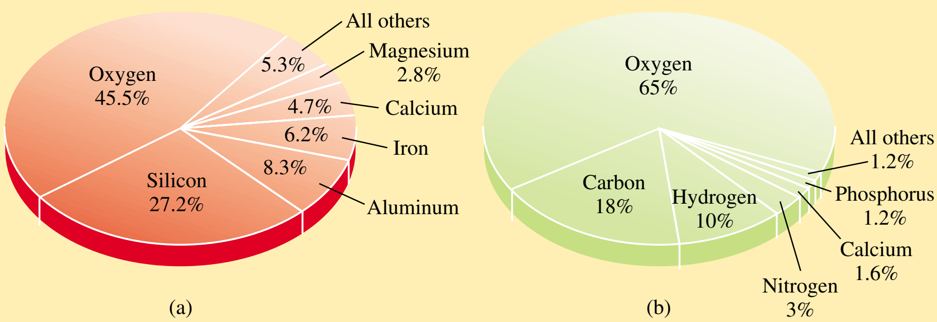
Structure of Earth's interior.

Essential Elements in the Human Body

ELEMENT	PERCENT BY MASS*	ELEMENT	PERCENT BY MASS*
Oxygen	65	Sodium	0.1
Carbon	18	Magnesium	0.05
Hydrogen	10	Iron	<0.05
Nitrogen	3	Cobalt	<0.05
Calcium	1.5	Copper	<0.05
Phosphorus	1.2	Zinc	<0.05
Potassium	0.2	Iodine	<0.05
Sulfur	0.2	Selenium	<0.01
Chlorine	0.2	Fluorine	<0.01

**Percent by mass* gives the mass of the element in grams present in a 100-g sample.

chemical names. Familiar examples are listed in Table 2.7.



(a) Natural abundance of the elements in percent by mass. For example, oxygen's abundance is 45.5 percent. This means that in a 100-g sample of Earth's crust there are, on the

3. An atom consists of a very dense central nucleus containing protons and neutrons, with electrons moving about the nucleus at a relatively large distance from it.
4. Protons are positively charged, neutrons have no charge, and electrons are negatively charged. Protons and neutrons have roughly the same mass, which is about 1840 times greater than the mass of an electron.
5. The atomic number of an element is the number of protons in the nucleus of an atom of the element; it determines the identity of an element. The mass number is the sum of the number of protons and the number of neutrons in the nucleus.
6. Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons.
7. Chemical formulas combine the symbols for the constituent elements with whole-number subscripts to show the type and number of atoms contained in the smallest unit of a compound.
8. The molecular formula conveys the specific number and type of atoms combined in each molecule of a compound. The empirical formula shows the simplest ratios of the atoms combined in a molecule.
9. Chemical compounds are either molecular compounds (in which the smallest units are discrete, individual molecules) or ionic compounds (in which positive and negative ions are held together by mutual attraction). Ionic compounds are made up of cations and anions, formed when atoms lose and gain electrons, respectively.
10. The names of many inorganic compounds can be deduced from a set of simple rules. The formulas can be written from the names of the compounds.

KEY WORDS

Acid, p. 58
Alkali metals, p. 46
Alkaline earth metals, p. 46
Allotrope, p. 50
Alpha (α) particles, p. 41
Alpha (α) rays, p. 41
Anion, p. 48
Atom, p. 39
Atomic number (Z), p. 45
Base, p. 60
Beta (β) particles, p. 41
Beta (β) rays, p. 41
Binary compound, p. 54
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Chemical formula, p. 50
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Molecule, p. 48
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Periodic table, p. 46
Polyatomic ion, p. 49
Polyatomic molecule, p. 48
Proton, p. 43
Radiation, p. 39
Radioactivity, p. 41
Ternary compound, p. 54

QUESTIONS AND PROBLEMS

STRUCTURE OF THE ATOM

Review Questions

2.1 Define the following terms: (a) α particle, (b) β particle, (c) γ ray, (d) X ray.
2.2 Name the types of radiation known to be emitted by radioactive elements.
2.3 Compare the properties of the following: α particles, cathode rays, protons, neutrons, electrons.
2.4 What is meant by the term "fundamental particle"?
2.5 Describe the contributions of the following scientists to our knowledge of atomic structure: J. J. Thomson, R. A. Millikan, Ernest Rutherford, James Chadwick.
2.6 Describe the experimental basis for believing that the nucleus occupies a very small fraction of the volume of the atom.

Problems

2.7 The diameter of a neutral helium atom is about 1×10^2 pm. Suppose that we could line up helium atoms side by side in contact with one another. Approximately how many atoms would it take to make the distance from end to end 1 cm?

2.8 Roughly speaking, the radius of an atom is about 10,000 times greater than that of its nucleus. If an atom were magnified so that the radius of its nucleus became 2.0 cm, about the size of a marble, what would be the radius of the atom in miles? (1 mi = 1609 m.)

ATOMIC NUMBER, MASS NUMBER, AND ISOTOPES

Review Questions

2.9 Use the helium-4 isotope to define atomic number and mass number. Why does a knowledge of atomic number enable us to deduce the number of electrons present in an atom?

2.10 Why do all atoms of an element have the same atomic number, although they may have different mass numbers?

2.11 What do we call atoms of the same elements with different mass numbers?

2.12 Explain the meaning of each term in the symbol $^{A}_{Z}X$.

Problems

2.13 What is the mass number of an iron atom that has 28 neutrons?

2.14 Calculate the number of neutrons of Pu-239.

2.15 For each of the following species, determine the number of protons and the number of neutrons in the nucleus:

$$^{3}_{2}He,\ ^{4}_{2}He,\ ^{24}_{12}Mg,\ ^{25}_{12}Mg,\ ^{48}_{22}Ti,\ ^{79}_{35}Br,\ ^{195}_{78}Pt$$

2.16 Indicate the number of protons, neutrons, and electrons in each of the following species:

$$^{15}_{7}N,\ ^{33}_{16}S,\ ^{63}_{29}Cu,\ ^{84}_{38}Sr,\ ^{130}_{56}Ba,\ ^{186}_{74}W,\ ^{202}_{80}Hg$$

2.17 Write the appropriate symbol for each of the following isotopes: (a) $Z = 11$, $A = 23$; (b) $Z = 28$, $A = 64$.

2.18 Write the appropriate symbol for each of the following isotopes: (a) $Z = 74$, $A = 186$; (b) $Z = 80$, $A = 201$.

THE PERIODIC TABLE

Review Questions

2.19 What is the periodic table, and what is its significance in the study of chemistry?

2.20 State two differences between a metal and a nonmetal.

2.21 Write the names and symbols for four elements in each of the following categories: (a) nonmetal, (b) metal, (c) metalloid.

2.22 Define, with two examples, the following terms: (a) alkali metals, (b) alkaline earth metals, (c) halogens, (d) noble gases.

Problems

2.23 Elements whose names end with *ium* are usually metals; sodium is one example. Identify a nonmetal whose name also ends with *ium*.

2.24 Describe the changes in properties (from metals to nonmetals or from nonmetals to metals) as we move (a) down a periodic group and (b) across the periodic table.

2.25 Consult a handbook of chemical and physical data (ask your instructor where you can locate a copy of the handbook) to find (a) two metals less dense than water, (b) two metals more dense than mercury, (c) the densest known solid metallic element, (d) the densest known solid nonmetallic element.

2.26 Group the following elements in pairs that you would expect to show similar chemical properties: K, F, P, Na, Cl, and N.

MOLECULES AND IONS

Review Questions

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

2.28 What are allotropes? Give an example. How are allotropes different from isotopes?

2.29 Describe the two commonly used molecular models.

2.30 Give an example of each of the following: (a) a monatomic cation, (b) a monatomic anion, (c) a polyatomic cation, (d) a polyatomic anion.

Problems

2.31 Identify the following as elements or compounds: NH_3, N_2, S_8, NO, CO, CO_2, H_2, SO_2.

2.32 Give two examples of each of the following: (a) a diatomic molecule containing atoms of the same element, (b) a diatomic molecule containing atoms of different elements, (c) a polyatomic molecule containing atoms of the same element, (d) a polyatomic molecule containing atoms of different elements.

2.33 Give the number of protons and electrons in each of the following common ions: Na^+, Ca^{2+}, Al^{3+}, Fe^{2+}, I^-, F^-, S^{2-}, O^{2-}, and N^{3-}.

2.34 Give the number of protons and electrons in each of the following common ions: K^+, Mg^{2+}, Fe^{3+}, Br^-, Mn^{2+}, C^{4-}, Cu^{2+}.

CHEMICAL FORMULAS

Review Questions

2.35 What does a chemical formula represent? What is the ratio of the atoms in the following molecular formulas? (a) NO, (b) NCl_3, (c) N_2O_4, (d) P_4O_6

2.36 Define molecular formula and empirical formula. What are the similarities and differences between the empirical formula and molecular formula of a compound?

2.37 Give an example of a case in which two molecules have different molecular formulas but the same empirical formula.

2.38 What does P_4 signify? How does it differ from 4P?

2.39 What is an ionic compound? How is electrical neutrality maintained in an ionic compound?

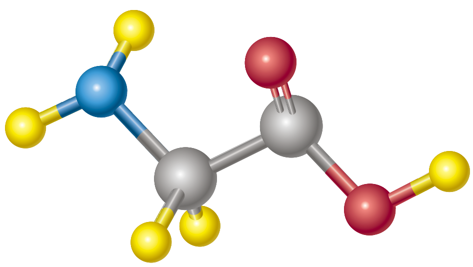
2.40 Explain why the chemical formulas of ionic compounds are always the same as their empirical formulas.

Problems

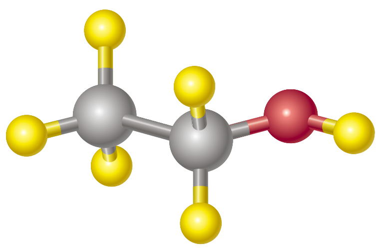
2.41 What are the empirical formulas of the following compounds? (a) C_2N_2, (b) C_6H_6, (c) C_9H_{20}, (d) P_4O_{10}, (e) B_2H_6

2.42 What are the empirical formulas of the following compounds? (a) Al_2Br_6, (b) $Na_2S_2O_4$, (c) N_2O_5, (d) $K_2Cr_2O_7$

2.43 Write the molecular formula of glycine, an amino acid present in proteins. The color codes are: gray (carbon), blue (nitrogen), red (oxygen), and yellow (hydrogen).



2.44 Write the molecular formula of ethanol. The color codes are: gray (carbon), red (oxygen), and yellow (hydrogen).



2.45 Which of the following compounds are likely to be ionic? Which are likely to be molecular? $SiCl_4$, LiF, $BaCl_2$, B_2H_6, KCl, C_2H_4

2.46 Which of the following compounds are likely to be ionic? Which are likely to be molecular? CH_4, NaBr, BaF_2, CCl_4, ICl, CsCl, NF_3

NAMING INORGANIC COMPOUNDS

Review Questions

2.47 What is the difference between inorganic compounds and organic compounds?

2.48 What are the four major categories of inorganic compounds?

2.49 Give an example each for a binary compound and a ternary compound.

2.50 What is the Stock system? What are its advantages over the older system of naming cations?

2.51 Explain why the formula HCl can represent two different chemical systems.

2.52 Define acids, bases, oxoacids, oxoanions, and hydrates.

Problems

2.53 Name the following compounds: (a) KH_2PO_4, (b) K_2HPO_4, (c) HBr (gas), (d) HBr (in water), (e) Li_2CO_3, (f) $K_2Cr_2O_7$, (g) NH_4NO_2, (h) HIO_3, (i) PF_5, (j) P_4O_6, (k) CdI_2, (l) $SrSO_4$, (m) $Al(OH)_3$.

2.54 Name the following compounds: (a) KClO, (b) Ag_2CO_3, (c) HNO_2, (d) $KMnO_4$, (e) $CsClO_3$, (f) KNH_4SO_4, (g) FeO, (h) Fe_2O_3, (i) $TiCl_4$, (j) NaH, (k) Li_3N, (l) Na_2O, (m) Na_2O_2.

2.55 Write the formulas for the following compounds: (a) rubidium nitrite, (b) potassium sulfide, (c) sodium hydrogen sulfide, (d) magnesium phosphate, (e) calcium hydrogen phosphate, (f) potassium dihydrogen phosphate, (g) iodine heptafluoride, (h) ammonium sulfate, (i) silver perchlorate, (j) boron trichloride.

2.56 Write the formulas for the following compounds: (a) copper(I) cyanide, (b) strontium chlorite, (c) perbromic acid, (d) hydroiodic acid, (e) disodium ammonium phosphate, (f) lead(II) carbonate, (g) tin(II) fluoride, (h) tetraphosphorus decasulfide, (i) mercury(II) oxide, (j) mercury(I) iodide, (k) selenium hexafluoride.

ADDITIONAL PROBLEMS

2.57 A sample of an uranium compound is found to be losing mass gradually. Explain what is happening to the sample.

2.58 In which one of the following pairs do the two species resemble each other most closely in chemical properties? Explain. (a) $^{1}_{1}H$ and $^{1}_{1}H^+$, (b) $^{14}_{7}N$ and $^{14}_{7}N^{3-}$, (c) $^{12}_{6}C$ and $^{13}_{6}C$

2.59 One isotope of a metallic element has mass number 65 and 35 neutrons in the nucleus. The cation derived from the isotope has 28 electrons. Write the symbol for this cation.

2.60 One isotope of a nonmetallic element has mass number 127 and 74 neutrons in the nucleus. The anion derived from the isotope has 54 electrons. Write the symbol for this anion.

2.61 The table below gives numbers of electrons, protons, and neutrons in atoms or ions of a number of elements. Answer the following: (a) Which of the species are neutral? (b) Which are negatively charged? (c) Which are positively charged? (d) What are the conventional symbols for all the species?

ATOM OR ION OF ELEMENT	A	B	C	D	E	F	G
Number of electrons	5	10	18	28	36	5	9
Number of protons	5	7	19	30	35	5	9
Number of neutrons	5	7	20	36	46	6	10

2.62 What is wrong with or ambiguous about the phrase "four molecules of NaCl"?

2.63 The following phosphorus sulfides are known: P_4S_3, P_4S_7, and P_4S_{10}. Do these compounds obey the law of multiple proportions?

2.64 Which of the following are elements, which are molecules but not compounds, which are compounds but not molecules, and which are both compounds and molecules? (a) SO_2, (b) S_8, (c) Cs, (d) N_2O_5, (e) O, (f) O_2, (g) O_3, (h) CH_4, (i) KBr, (j) S, (k) P_4, (l) LiF

2.65 Why is magnesium chloride ($MgCl_2$) not called magnesium(II) chloride?

2.66 Some compounds are better known by their common names than by their systematic chemical names. Give the chemical formulas of the following substances: (a) dry ice, (b) table salt, (c) laughing gas, (d) marble (chalk, limestone), (e) quicklime, (f) slaked lime, (g) baking soda, (h) washing soda, (i) gypsum, (j) milk of magnesia.

2.67 Fill in the blanks in the following table:

SYMBOL		$^{54}_{26}Fe^{2+}$			
PROTONS	5			79	86
NEUTRONS	6		16	117	136
ELECTRONS	5		18	79	
NET CHARGE			−3		0

2.68 (a) Which elements are most likely to form ionic compounds? (b) Which metallic elements are most likely to form cations with different charges?

2.69 What ion is each of the following most likely to form in ionic compounds: (a) Li, (b) S, (c) I, (d) N, (e) Al, (f) Cs, (g) Mg?

2.70 Which of the following symbols provides more information about the atom: ^{23}Na or $_{11}Na$? Explain.

2.71 Write the chemical formulas and names of acids that contain Group 7A elements. Do the same for elements in Groups 3A, 4A, 5A, and 6A.

2.72 Of the 112 elements known, only two are liquids at room temperature (25°C). What are they? (*Hint:* One element is a familiar metal and the other element is in Group 7A.)

2.73 For the noble gases (the Group 8A elements), $^{4}_{2}He$, $^{20}_{10}Ne$, $^{40}_{18}Ar$, $^{84}_{36}Kr$, and $^{132}_{54}Xe$, (a) determine the number of protons and neutrons in the nucleus of each atom, and (b) determine the ratio of neutrons to protons in the nucleus of each atom. Describe any general trend you discover in the way this ratio changes with increasing atomic number.

2.74 List the elements that exist as gases at room temperature. (*Hint:* These elements can be found in Groups 5A, 6A, 7A, and 8A.)

2.75 The Group 1B metals, Cu, Ag, and Au, are called coinage metals. What chemical properties make them specially suitable for making coins and jewels?

2.76 The elements in Group 8A of the periodic table are called noble gases. Can you suggest what "noble" means in this context?

2.77 The formula for calcium oxide is CaO. What are the formulas for magnesium oxide and strontium oxide?

2.78 A common mineral of barium is barytes, or barium sulfate ($BaSO_4$). Because elements in the same periodic group have similar chemical properties, we might expect to find some radium sulfate ($RaSO_4$) mixed with barytes since radium is the last member of Group 2A. However, the only source of radium compounds in nature is in uranium minerals. Why?

2.79 List five elements each that are (a) named after places, (b) named after people, (c) named after a color. (*Hint:* See Appendix 1.)

2.80 Name the only country that is named after an element. (*Hint:* This country is in South America.)

2.81 Fluorine reacts with hydrogen (H) and deuterium (D) to form hydrogen fluoride (HF) and deuterium fluoride (DF), where deuterium ($^{2}_{1}H$) is an isotope of hydrogen. Would a given amount of fluorine react with different masses of the two hydrogen isotopes? Does this violate the law of definite proportion? Explain.

2.82 Predict the formula and name of a binary compound formed from the following elements: (a) Na and H, (b) B and O, (c) Na and S, (d) Al and F, (e) F and O, (f) Sr and Cl.

2.83 Identify each of the following elements: (a) a halogen whose anion contains 36 electrons, (b) a radioactive noble gas with 86 protons, (c) a Group 6A element whose anion contains 36 electrons, (d) An alkali metal cation that contains 36 electrons, (e) a Group 4A cation that contains 80 electrons.

Answers to Practice Exercises: **2.1** 29 protons, 34 neutrons, and 29 electrons. **2.2** $CHCl_3$. **2.3** $C_4H_5N_2O$. **2.4** (a) Lead(II) oxide, (b) lithium sulfite. **2.5** (a) Rb_2SO_4, (b) BaH_2. **2.6** (a) Nitrogen trifluoride, (b) dichlorine heptoxide. **2.7** (a) SF_4, (b) N_2O_5. **2.8** (a) Hypobromous acid, (b) hydrogen sulfate ion.