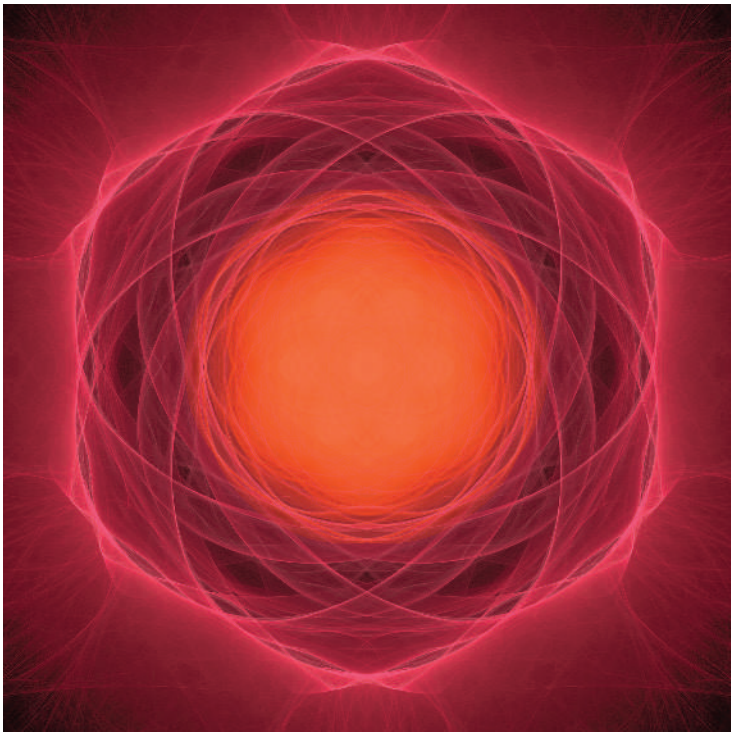
2 Atoms, Molecules, and Ions

An artist's view of an atom.


Courtesy RochVanh

Atom from atom yawns as far
As moon from earth, or star
from star.

—RALPH WALDO EMERSON
Atoms

Chapter Outline

To learn chemistry, you must become familiar with the building blocks that chemists use to describe the structure of matter. These include

- *atoms* (Section 2-1), composed of electrons, protons, and neutrons (Section 2-2) and their quantitative properties—atomic mass and atomic number (Section 2-3).
- *molecules*, the building blocks of several elements and a great many compounds. Molecular substances can be identified by their formulas (Section 2-5) or their names (Section 2-7)
- *ions*, species of opposite charge found in all ionic compounds. Using relatively simple principles, it is possible to derive the formulas (Section 2-6) and names (Section 2-7) of ionic compounds.

Early in this chapter (Section 2-4), we will introduce a classification system for elements known as the *periodic table*. It will prove useful in this chapter and throughout the remainder of the text.

2-1 Atoms and the Atomic Theory

In 1808, an English scientist and schoolteacher, John Dalton, developed the atomic model of matter that underlies modern chemistry. Three of the main postulates of modern **atomic theory**, all of which Dalton suggested in a somewhat different form, are stated below and illustrated in Figure 2.1.

1. *An element is composed of tiny particles called atoms.* All atoms of a given element have the same chemical properties. Atoms of different elements show different properties.
2. *In an ordinary chemical reaction, atoms move from one substance to another, but no atom of any element disappears or is changed into an atom of another element.*
3. *Compounds are formed when atoms of two or more elements combine.* In a given compound, the relative numbers of atoms of each kind are definite and constant. In general, these relative numbers can be expressed as integers or simple fractions.

On the basis of Dalton's theory, the **atom** can be defined as the smallest particle of an element that can enter into a chemical reaction.

2-2 Components of the Atom

Like any useful scientific theory, the atomic theory raised more questions than it answered. Scientists wondered whether atoms, tiny as they are, could be broken down into still smaller particles. Nearly 100 years passed before the existence of subatomic particles was confirmed by experiment. Two future Nobel laureates did pioneer work in this area. J. J. Thomson was an English physicist working at the Cavendish Laboratory at Cambridge. Ernest Rutherford, at one time a student of Thomson's (Figure 2.2), was a native of New Zealand. Rutherford carried out his research at McGill University in Montreal and at Manchester and Cambridge in England. He was clearly the greatest experimental physicist of his time, and one of the greatest of all time.

Electrons

The first evidence for the existence of subatomic particles came from studies of the conduction of electricity through gases at low pressures. When the glass tube shown in Figure 2.3 is partially evacuated and connected to a spark coil,

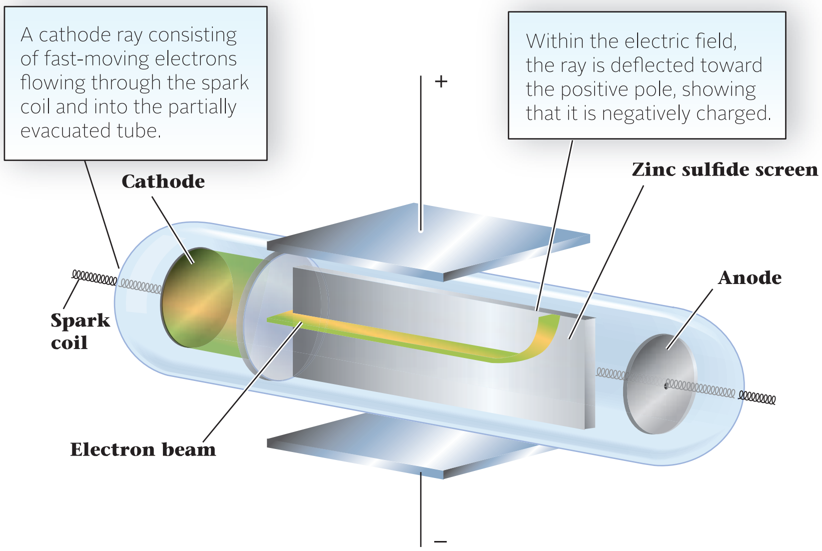


Figure 2.3 A cathode ray passing through an electric field.

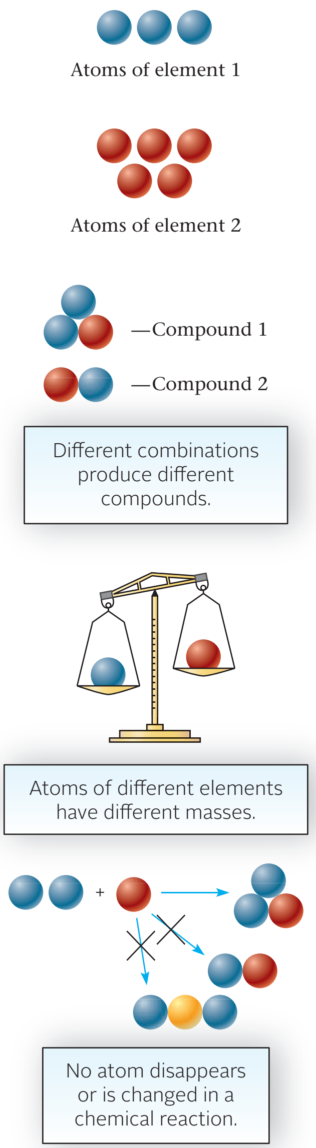


Figure 2.1 Some features of Dalton's atomic theory.

Atoms are indeed tiny (they have diameters of about 10^{-10} m).

Figure 2.2 J. J. Thomson and Ernest Rutherford *(right)*. They are talking, perhaps about nuclear physics, but more likely about yesterday's cricket match.

CHEMISTRY THE **HUMAN** SIDE

John Dalton was a quiet, unassuming man and a devout Quaker. When presented to King William IV of England, Dalton refused to wear the colorful court robes because of his religion. His friends persuaded him to wear the scarlet robes of Oxford University, from which he had a doctor's degree. Dalton was color-blind, so he saw himself clothed in gray.

Dalton was a prolific scientist who made contributions to biology and physics as well as chemistry. At a college in Manchester, England, he did research and spent as many as 20 hours a week lecturing in mathematics and the physical sciences. Dalton never married; he said once, "My head is too full of triangles, chemical properties, and electrical experiments to think much of marriage."

Dalton's atomic theory explained three of the basic laws of chemistry:

The *law of conservation of mass:* This states that *there is no detectable change in mass in an ordinary chemical reaction.* If atoms are conserved in a reaction (postulate 2 of the atomic theory), mass will also be conserved.

The *law of constant composition:* This tells us that *a compound always contains the same elements in the same proportions by mass.* If the atom ratio of the elements in a compound is fixed (postulate 3), their proportions by mass must also be fixed.

The Edgar Fahs Smith Memorial Collection in the History of Chemistry, Department of Special Collections, Van Pelt-Dietrich Library, University of Pennsylvania

John Dalton
(1766–1844)

The *law of multiple proportions:* This law, formulated by Dalton himself, was crucial to establishing atomic theory. It applies to situations in which two elements form more than one compound. The law states that in these compounds, *the masses of one element that combine with a fixed mass of the second element are in a ratio of small whole numbers.*

The validity of this law depends on the fact that atoms combine in simple, whole-number ratios (postulate 3). Its relation to atomic theory is further illustrated in Figure A.

Mama G. Clarke

Figure A Chromium-oxygen compounds and the law of multiple proportions. Chromium forms two different compounds with oxygen, as shown by their different colors. In the green compound there are two chromium atoms for every three oxygen atoms (2Cr : 3O) and 2.167 g of chromium per gram of oxygen. In the red compound there is one chromium atom for every three oxygen atoms (1Cr : 3O) and 1.083 g of chromium per gram of oxygen. The ratio of the chromium masses, 2.167 : 1.083, is that of two small whole numbers, 2.167 : 1.083 = 2 : 1, an illustration of the law of multiple proportions.

an electric current flows through it. Associated with this flow are colored rays of light called *cathode rays,* which are bent by both electric and magnetic fields. From a careful study of this deflection, J. J. Thomson showed in 1897 that the rays consist of a stream of negatively charged particles, which he called **electrons.** We now know that electrons are common to all atoms, carry a unit negative charge (−1), and have a very small mass, roughly 1/2000 that of the lightest atom.

Every atom contains a definite number of electrons. This number, which runs from 1 to more than 100, is characteristic of a neutral atom of a particular element. All atoms of hydrogen contain one electron; all atoms of the element uranium contain 92 electrons. We will have more to say in Chapter 6 about how these electrons are arranged relative to one another. Right now, you need only know that they are found in the outer regions of the atom, where they form what amounts to a cloud of negative charge.

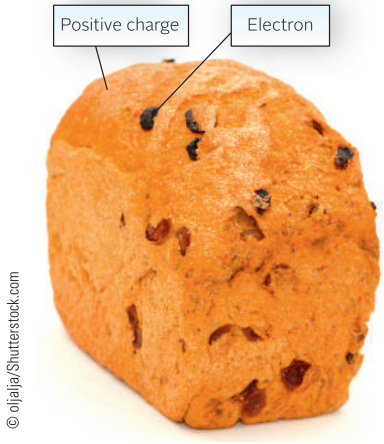


Figure 2.4 Representation of Thomson's model. The raisins are representative of the electrons distributed according to Thomson's model.

Protons and Neutrons; the Atomic Nucleus

After J. J. Thomson discovered the negatively charged particles (electrons) in the atom, he proposed the structure of the atom to consist of a positively charged sphere with the negatively charged electrons embedded in that sphere. It was known at the time to be the "plum pudding model." Today, we would probably call it the "raisin bread model," where the raisins are the electrons distributed in the positively charged nucleus, the bread (Figure 2.4).

A series of experiments carried out under the direction of Ernest Rutherford in 1911 shaped our ideas about the nature of the atom. He and his students bombarded a piece of thin gold foil (Figure 2.5, p. 26) with α particles (helium atoms minus their electrons). With a fluorescent screen, they observed the extent to which the α particles were scattered. Most of the particles went through the foil unchanged in direction; a few, however, were reflected back at acute angles. This was a totally unexpected result, inconsistent with the model of the atom in vogue at that time. In Rutherford's words, "It was as though you had fired a 15-inch shell at a piece of tissue paper and it had bounced back and hit you." By a mathematical analysis of the forces involved, Rutherford showed that the scattering was caused by a small, positively charged **nucleus** at the center of the gold atom. Most of the atom is empty space, which explains why most of the bombarding particles passed through the gold foil undeflected.

i If the nucleus were the size of your head, the electron would be 5 miles away.

Since Rutherford's time scientists have learned a great deal about the properties of atomic nuclei. For our purposes in chemistry, the nucleus of an atom can be considered to consist of two different types of particles (Table 2.1):

1. The **proton,** which has a mass nearly equal to that of an ordinary hydrogen atom. The proton carries a unit positive charge (+1), equal in magnitude to that of the electron (−1).
2. The **neutron,** an uncharged particle with a mass slightly greater than that of a proton.

Because protons and neutrons are much heavier than electrons, most of the mass of an atom (>99.9%) is concentrated in the nucleus, even though the volume of the nucleus is much smaller than that of the atom.

i The diameter of an atom is 10,000 times that of its nucleus.

Table 2.1 Properties of Subatomic Particles

Particle	Location	Relative Charge	Relative Mass*
Proton	Nucleus	+1	1.00728
Neutron	Nucleus	0	1.00867
Electron	Outside nucleus	−1	0.00055

*These are expressed in atomic mass units.

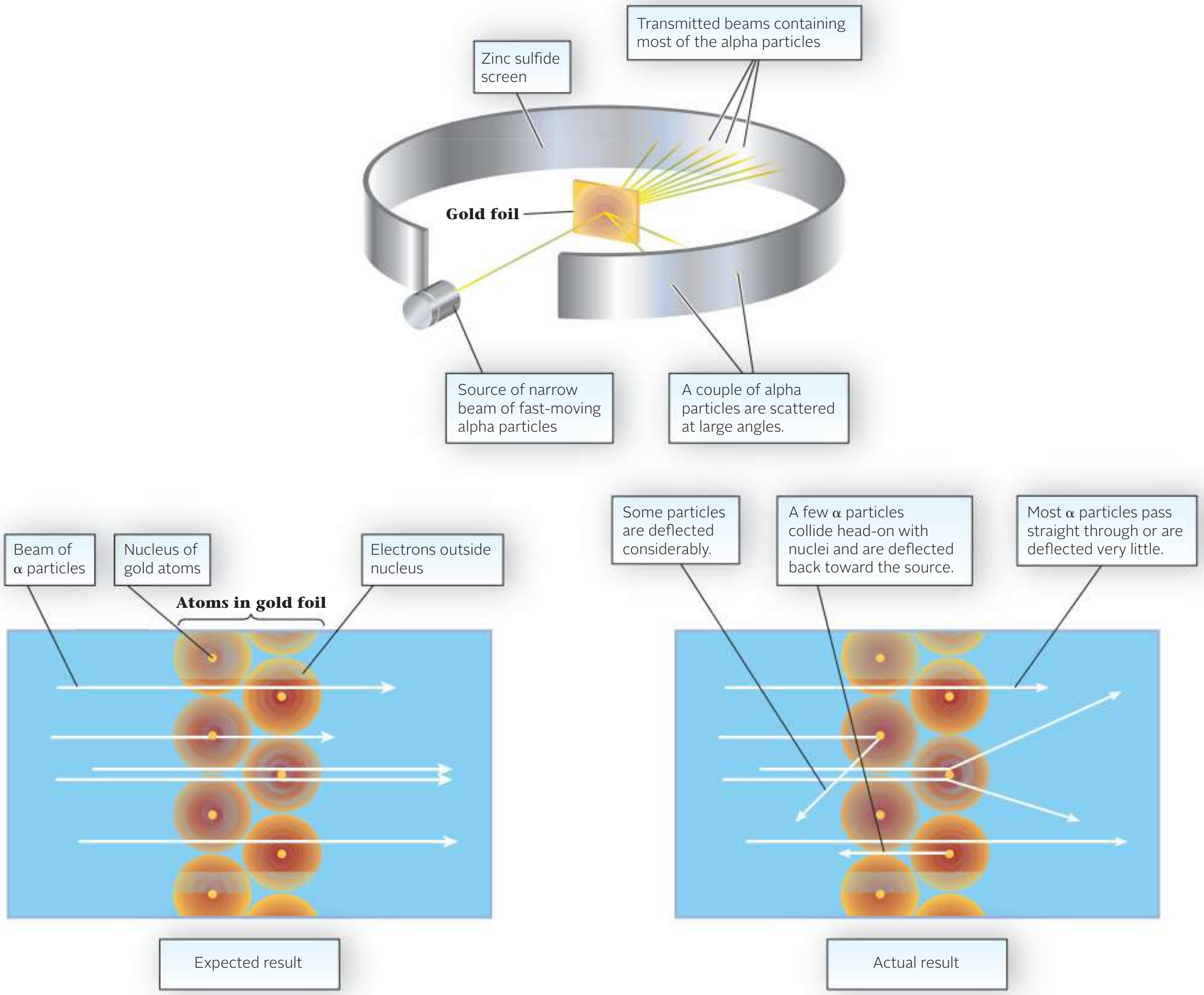


Figure 2.5 **Rutherford's α particle scattering experiment.**

2-3 Quantitative Properties of the Atom

Atomic Number

All the atoms of a particular element have the same number of protons in the nucleus. This number is a basic property of an element, called its **atomic number** and given the symbol Z:

$$Z = \text{number of protons}$$

In a neutral atom, the number of protons in the nucleus is exactly equal to the number of electrons outside the nucleus. Consider, for example, the elements hydrogen ($Z = 1$) and uranium ($Z = 92$). All hydrogen atoms have one proton in the nucleus; all uranium atoms have 92. In a neutral hydrogen atom there is one electron outside the nucleus; in a uranium atom there are 92.

H atom:	1 proton, 1 electron	$Z = 1$
U atom:	92 protons, 92 electrons	$Z = 92$

Mass Numbers; Isotopes

The **mass number** of an atom, given the symbol A, is found by adding up the number of protons and neutrons in the nucleus:

$$A = \text{number of protons} + \text{number of neutrons}$$

All atoms of a given element have the same number of protons, hence the same atomic number. They may, however, differ from one another in mass and therefore in mass number. This can happen because, although the number of protons in an atom of an element is fixed, the number of neutrons is not. It may vary and often does. Consider the element hydrogen ($Z = 1$). There are three different kinds of hydrogen atoms. They all have one proton in the nucleus. A light hydrogen atom (the most common type) has no neutrons in the nucleus ($A = 1$). Another type of hydrogen atom (deuterium) has one neutron ($A = 2$). Still a third type (tritium) has two neutrons ($A = 3$).

Number of neutrons $= A - Z$

An atom that contains the same number of protons but a different number of neutrons is called an **isotope**. The three kinds of hydrogen atoms just described are isotopes of that element. They have masses that are very nearly in the ratio 1:2:3. Among the isotopes of the element uranium are the following:

Isotope	Z	A	Number of Protons	Number of Neutrons
Uranium-235	92	235	92	143
Uranium-238	92	238	92	146

© Cengage Learning/Charles D. Winters

$^2_1H_2O$ ice and $^1_1H_2O$ ice. Solid deuterium oxide (*bottom*) is more dense than liquid water and sinks, whereas the ordinary water ice (*top*) is less dense than liquid water and floats.

The composition of a nucleus is shown by its **nuclear symbol**. Here, the atomic number appears as a subscript at the lower left of the symbol of the element. The mass number is written as a superscript at the upper left.

$$\begin{matrix}\text{Mass number} \longrightarrow \\ \text{Atomic number} \longrightarrow\end{matrix} {}^{A}_{Z}\text{X} \longleftarrow \text{element symbol}$$

The nuclear symbols for the isotopes of hydrogen and uranium referred to above are

$$^1_1\text{H},\ ^2_1\text{H},\ ^3_1\text{H} \qquad ^{235}_{92}\text{U},\ ^{238}_{92}\text{U}$$

Quite often, isotopes of an element are distinguished from one another by writing the mass number after the symbol of the element. The isotopes of uranium are often referred to as U-235 and U-238.

EXAMPLE 2.1

a An isotope of antimony (Sb, $Z = 51$) is used as a tracer to detect leaks in oil pipelines. This isotope has 73 neutrons in its nucleus. What is its nuclear symbol?

ANALYSIS	
Information given:	Z (51); number of neutrons, n (73)
Asked for:	nuclear symbol

STRATEGY

1. Note that Z stands for the atomic number or the number of protons p^+.
2. Recall that a nuclear symbol is written ^A_ZX where A stands for the number of neutrons (n) and protons (p^+).
3. Use this strategy for parts (b) and (c) also.

continued

SOLUTION	
nuclear symbol	$Z = p^+ = 51$; $A = p^+ + n = 51 + 73 = 124$; ${}^{A}_{Z}\text{Sb} = {}^{124}_{51}\text{Sb}$
b One of the most harmful components of nuclear waste is a radioactive isotope of strontium, ${}^{90}_{38}\text{Sr}$; it can be deposited in your bones, where it replaces calcium. How many protons are there in the nucleus of Sr-90? How many neutrons?	
ANALYSIS	
Information given:	nuclear symbol: ${}^{90}_{38}\text{Sr}$
Asked for:	p^+; n
SOLUTION	
protons	$Z = p^+ = 38$
neutrons	$A = p^+ + n = 90$; $90 = 38 + n$; $n = 90 - 38 = 52$
c Write the symbol for the element used as a medical tool for scanning lungs. It has 54 protons and 79 neutrons.	
ANALYSIS	
Information given:	$p^+ = 54$; $n = 79$
Information implied:	identity of the element
Asked for:	nuclear symbol
SOLUTION	
nuclear symbol	$Z = p^+ = 54$ (placed on bottom left of the element) $A = p^+ + n = 54 + 79 = 133$ The element (X) is xenon identified by its atomic number Z. nuclear symbol: ${}^{A}_{Z}\text{X} = {}^{133}_{54}\text{Xe}$

Atomic Masses

Individual atoms are far too small to be weighed on a balance. However, as you will soon see, it is possible to determine quite accurately the relative masses of different atoms and molecules. Indeed, it is possible to go a step further and calculate the actual masses of these tiny building blocks of matter.

Relative masses of atoms of different elements are expressed in terms of their *atomic masses* (often referred to as *atomic weights*). The **atomic mass** of an element indicates how heavy, on the average, one atom of that element is compared with an atom of another element.

To set up a scale of atomic masses, it is necessary to establish a standard value for one particular species. The modern atomic mass scale (sometimes called the **Carbon-12 scale**) is based on the most common isotope of carbon, ${}^{12}_{6}\text{C}$. This isotope is assigned a mass of exactly 12 **atomic mass units (amu):**

1 amu = 1/12 mass C atom ≈ mass H atom.

$$\text{mass of C-12 atom} = 12 \text{ amu (exactly)}$$

It follows that an atom half as heavy as a C-12 atom would weigh 6 amu, an atom twice as heavy as C-12 would have a mass of 24 amu, and so on.

Isotopic Abundances

Relative masses of individual atoms can be determined using a mass spectrometer (Figure 2.6). Gaseous atoms or molecules at very low pressures are ionized by removing one or more electrons. The cations formed are accelerated by a potential

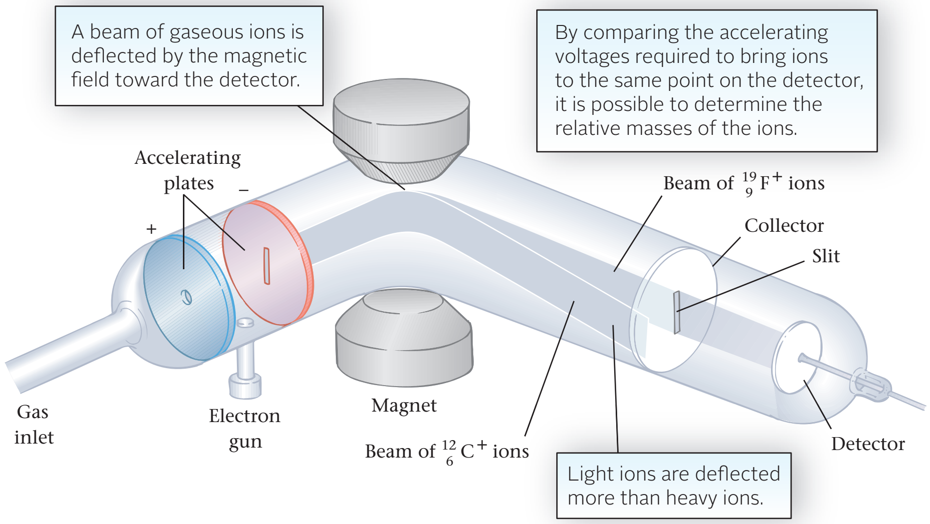


Figure 2.6 The mass spectrometer.

of 500 to 2000 V toward a magnetic field, which deflects the ions from their straight-line path. The extent of deflection is inversely related to the mass of the ion. By measuring the voltages required to bring two ions of different mass to the same point on the detector, it is possible to determine their relative masses. For example, using a mass spectrometer, it is found that a $^{19}_{9}F$ atom is 1.583 times as heavy as a $^{12}_{6}C$ atom and so has a mass of

$$1.583 \times 12.00 \text{ amu} = 19.00 \text{ amu}$$

As it happens, naturally occurring fluorine consists of a single isotope, $^{19}_{9}F$. It follows that the atomic mass of the element fluorine must be the same as that of F-19, 19.00 amu. The situation with most elements is more complex because they occur in nature as a mixture of two or more isotopes. To determine the atomic mass of such an element, it is necessary to know not only the masses of the individual isotopes but also their atom percents (*isotopic abundances*) in nature.

Fortunately, isotopic abundances as well as isotopic masses can be determined by mass spectrometry. The situation with chlorine, which has two stable isotopes, Cl-35 and Cl-37, is shown in Figure 2.7. The atomic masses of the two isotopes are determined in the usual way. The relative abundances of these isotopes are proportional to the heights of the recorder peaks or, more accurately, to the areas under these peaks. For chlorine, the data obtained from the mass spectrometer are

	Atomic Mass	Abundance
Cl-35	34.97 amu	75.53%
Cl-37	36.97 amu	24.47%

We interpret this to mean that, in elemental chlorine, 75.53% of the atoms have a mass of 34.97 amu, and the remaining atoms, 24.47% of the total, have a mass of 36.97 amu. With this information we can readily calculate the atomic mass of chlorine using the general equation

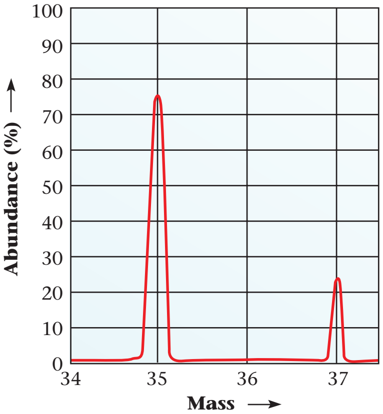


Figure 2.7 Mass spectrum of chlorine. Elemental chlorine (Cl_2) contains only two isotopes: 34.97 amu (75.53%) and 36.97 amu (24.47%).

atomic mass Y =

$$(\text{atomic mass } Y_1) \times \frac{\%\ Y_1}{100\%} + (\text{atomic mass } Y_2) \times \frac{\%\ Y_2}{100\%} + \cdots \qquad \textbf{(2.1)}$$

where $Y_1, Y_2, \ldots$ are isotopes of element Y.

$$\text{atomic mass Cl} = 34.97 \text{ amu} \times \frac{75.53}{100.0} + 36.97 \text{ amu} \times \frac{24.47}{100.0} = 35.46 \text{ amu}$$

Atomic masses calculated in this manner, using data obtained with a mass spectrometer, can in principle be precise to seven or eight significant figures. The accuracy of tabulated atomic masses is limited mostly by variations in natural abundances. Sulfur is an interesting case in point. It consists largely of two isotopes, $^{32}_{16}S$ and $^{34}_{16}S$. The abundance of sulfur-34 varies from about 4.18% in sulfur deposits in Texas and Louisiana to 4.34% in volcanic sulfur from Italy. This leads to an uncertainty of 0.006 amu in the atomic mass of sulfur.

If the atomic mass of an element is known *and* if it has only two stable isotopes, their abundances can be calculated from the general equation cited above.

EXAMPLE 2.2

Bromine is a red-orange liquid with an average atomic mass of 79.90 amu. Its name is derived from the Greek word *bromos* ($\beta\rho o\mu o\varsigma$), which means stench. It has two naturally occurring isotopes: Br-79 (78.92 amu) and Br-81 (80.92 amu). What is the abundance of the heavier isotope?

ANALYSIS

Information given:	Br-81 mass (80.92 amu); Br-79 mass (78.92 amu) average atomic mass (79.90)
Asked for:	abundance of Br-81

STRATEGY

1. All abundances must add up to 100%
2. Recall the formula relating abundance and atomic mass (Equation 2.1)

$$\text{atomic mass Y} = \left(\text{atomic mass } Y_1 \times \frac{\%\ Y_1}{100\%}\right) + \left(\text{atomic mass } Y_2 \times \frac{\%\ Y_2}{100\%}\right) + \cdots$$

SOLUTION

1. % abundances

Br-81: x; Br-79: $100 - x$

2. Substitute into Equation 2.1.

$$79.90 \text{ amu} = 78.92 \text{ amu}\left(\frac{100 - x}{100}\right) + 80.92 \text{ amu}\left(\frac{x}{100}\right)$$

3. Solve for x.

$$79.90 = 0.7892(100 - x) + 0.8092\,x$$

$$79.90 = 78.92 - 0.7892\,x + 0.8092\,x$$

$$x = 49\%$$

END POINT

The atomic mass of Br, 79.90, is just about halfway between the masses of the two isotopes, 78.92 and 80.92. So, it is reasonable that it should contain nearly equal amounts of the two isotopes.

The inside cover or opening pages of your text show the symbols of all the elements arranged in a particular way. This is called a **periodic table.** We will have more to say about this table in the next section (Section 2-4). For now, you can use it to find the element's average atomic mass (rounded to four digits and written below the element's symbol). Its atomic number, Z, is above the symbol.

The average atomic mass shown in the periodic table is not equal to the mass number.

Masses of Individual Atoms; Avogadro's Number

For most purposes in chemistry, it is sufficient to know the relative masses of different atoms. Sometimes, however, it is necessary to go one step further and calculate the mass in grams of individual atoms. Let us consider how this can be done.

To start with, consider the elements helium and hydrogen. A helium atom is about four times as heavy as a hydrogen atom (He = 4.003 amu, H = 1.008 amu). It follows that a sample containing 100 helium atoms weighs about four times as much as a sample containing 100 hydrogen atoms. Again, comparing samples of the two elements containing a million atoms each, the masses will be in a 4 (helium) to 1 (hydrogen) ratio. Turning this argument around, it follows that a sample of helium weighing four grams must contain very nearly the same number of atoms as a sample of hydrogen weighing one gram. More precisely

If a nickel weighs twice as much as a dime, there are equal numbers of coins in 1000 g of nickels and 500 g of dimes.

no. of He atoms in 4.003 g helium = no. of H atoms in 1.008 g hydrogen

This reasoning is readily extended to other elements. A sample of any element with a mass in grams equal to its atomic mass contains the same number of atoms, N_A, regardless of the identity of the element.

The question now arises as to the numerical value of N_A; that is, how many atoms are in 4.003 g of helium, 1.008 g of hydrogen, 32.07 g of sulfur, and so on? As it happens, this problem is one that has been studied for at least a century. Several ingenious experiments have been designed to determine this number, known as **Avogadro's number** and given the symbol N_A. As you can imagine, it is huge. (Remember that atoms are tiny. There must be a lot of them in 4.003 g of He, 1.008 g of H, and so on.) To four significant figures,

$$N_A = 6.022 \times 10^{23}$$

To get some idea of how large this number is, suppose the entire population of the world were assigned to counting the atoms in 4.003 g of helium. If each person counted one atom per second and worked a 48-hour week, the task would take more than 10 million years.

Most people have better things to do.

The importance of Avogadro's number in chemistry should be clear. ***It represents the number of atoms of an element in a sample whose mass in grams is numerically equal to the atomic mass of the element.*** Thus there are

6.022×10^{23} H atoms in 1.008 g H	atomic mass H = 1.008 amu
6.022×10^{23} He atoms in 4.003 g He	atomic mass He = 4.003 amu
6.022×10^{23} S atoms in 32.07 g S	atomic mass S = 32.07 amu

Knowing Avogadro's number and the atomic mass of an element, it is possible to calculate the mass of an individual atom (Example 2.3a). You can also determine the number of atoms in a weighed sample of any element (Example 2.3b).

Avogadro's number has been given a special name—mole. We will have more to say about the mole in Section 3-1.

EXAMPLE 2.3

Consider arsenic (As), a favorite poison used in crime stories. This element is discussed at the end of Chapter 1. Avogadro's number is 6.022×10^{23}.

a What is the mass of an arsenic atom?

ANALYSIS	
Information given:	Avogadro's number (6.022×10^{23})
Information implied:	atomic mass
Asked for:	mass of an arsenic atom

continued

STRATEGY

Change atoms to grams (atoms $\longrightarrow$ g) by using the conversion factor

$$\frac{6.022 \times 10^{23}\text{ atoms}}{\text{atomic mass}}$$

SOLUTION

mass of an As atom	$1\text{ atom As} \times \dfrac{74.92\text{ g As}}{6.022 \times 10^{23}\text{ atoms As}} = 1.244 \times 10^{-22}\text{ g}$

b How many atoms are there in ten grams of arsenic?

ANALYSIS

Information given:	mass of sample (10.00 g) from (a) mass of one As atom (1.244×10^{-22} g/atom)
Asked for:	number of atoms in a ten-gram sample

STRATEGY

Change grams to atoms (g $\longrightarrow$ atom) by using the conversion factor

$$\frac{1\text{ atom}}{1.244 \times 10^{-22}\text{ g}}$$

SOLUTION

atoms of As	$10.00\text{ g As} \times \dfrac{1\text{ atom As}}{1.244 \times 10^{-22}\text{ g As}} = 8.038 \times 10^{22}\text{ atoms As}$

c How many protons are there in 0.1500 lb of arsenic?

ANALYSIS

Information given:	mass of sample (0.1500 lbs) from (a) mass of one As atom (1.244×10^{-22} g/atom)
Information implied:	atomic number pounds to grams conversion factor
Asked for:	number of protons in 0.1500 lb As

STRATEGY

Change pounds to grams, grams to atoms, and atoms to protons by using the conversion factors

$$\frac{453.6\text{ g}}{1\text{ lb}} \qquad \frac{\text{no. of protons }(Z)}{1\text{ atom}} \qquad \frac{1\text{ atom}}{1.244 \times 10^{-22}\text{ g}}$$

and follow the plan:

lb $\longrightarrow$ g $\longrightarrow$ atom $\longrightarrow$ proton

SOLUTION

number of protons	$0.1500\text{ lb} \times \dfrac{453.6\text{ g}}{1\text{ lb}} \times \dfrac{1\text{ atom}}{1.244 \times 10^{-22}\text{ g}} \times \dfrac{33\text{ protons}}{1\text{ atom As}} = 1.805 \times 10^{25}\text{ protons}$

END POINT

Because atoms are so tiny, we expect their mass to be very small: 1.244×10^{-22} g sounds reasonable. Conversely, it takes a lot of atoms, in this case, 8.038×10^{22} atoms, to weigh ten grams.

2-4 Introduction to the Periodic Table

From a microscopic point of view, an element is a substance all of whose atoms have the same number of protons, that is, the same atomic number. The chemical properties of elements depend upon their atomic numbers, which can be read from the periodic table. A complete periodic table that lists symbols, atomic numbers, and atomic masses is given on the inside front cover or opening pages of this text. For our purposes in this chapter, the abbreviated table in Figure 2.8 will suffice.

Periods and Groups

The horizontal rows in the table are referred to as **periods**. The first period consists of the two elements hydrogen (H) and helium (He). The second period starts with lithium (Li) and ends with neon (Ne).

The vertical columns are known as **groups**. Historically, many different systems have been used to designate the different groups. Both Arabic and Roman numerals have been used in combination with the letters A and B. The system used in this text is the one recommended by the International Union of Pure and Applied Chemistry (IUPAC) in 1985. The groups are numbered from 1 to **18**, starting at the left.

Other periodic tables label the groups differently, but the elements have the same position.

Elements falling in Groups 1, 2, 1**3**, 1**4**, 1**5**, 1**6**, 1**7**, and 1**8*** are referred to as **main-group elements.** The ten elements in the center of each of periods 4 through 6 are called **transition metals**; they fall in Groups 3 through 12. The first transition series (period 4) starts with Sc (Group 3) and ends with Zn (Group 12).

You'll have to wait until Chapter 7 to learn why the second digit of some group numbers are in bold type.

The metals in Groups 1**3**, 1**4**, and 1**5**, which lie to the right of the transition metals (Ga, In, Tl, Sn, Pb, Bi), are often referred to as **post-transition metals**.

Certain main groups are given special names. The elements in Group 1, at the far left of the periodic table, are called **alkali metals**; those in Group 2 are

	1	2	3	4	5	6	7	8	9	10	11	12	1**3**	1**4**	1**5**	1**6**	1**7**	1**8**
1	1 H																1 H	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn

Metals Metalloids Nonmetals

Figure 2.8 Periodic table. The group numbers stand above the columns. The numbers at the left of the rows are the period numbers. The black line separates the metals from the nonmetals. (*Note:* A complete periodic table is given inside the cover or opening pages of your text.)

*Prior to 1985, Groups 1**3** to 1**8** were commonly numbered 3 to 8 or 3A to 8A in the United States.

© Cengage Learning/Charles D. Winters

Figure 2.9 Inert gases. Neon is used in advertising signs. The gas in tubes adds both color and light. The test kit is used to detect radon in home basements.

referred to as **alkaline earth metals**. As we move to the right, the elements in Group 17 are called **halogens**; at the far right, the **noble gases** (also called "unreactive gases") constitute Group 18.

Elements in the same main group show very similar chemical properties. For example,

- lithium (Li), sodium (Na), and potassium (K) in Group 1 all react vigorously with water to produce hydrogen gas.
- helium (He), neon (Ne) (Figure 2.9), and argon (Ar) in Group 18 do not react with any other substances.

On the basis of observations such as these, we can say that ***the periodic table is an arrangement of elements, in order of increasing atomic number, in horizontal rows of such a length that elements with similar chemical properties fall directly beneath one another in vertical groups.***

The person whose name is most closely associated with the periodic table is Dmitri Mendeleev (1834–1907), a Russian chemist. In writing a textbook of general chemistry, Mendeleev devoted separate chapters to families of elements with similar properties, including the alkali metals, the alkaline earth metals, and the halogens. Reflecting on the properties of these and other elements, he proposed in 1869 a primitive version of today's periodic table. Mendeleev shrewdly left empty spaces in his table for new elements yet to be discovered. Indeed, he predicted detailed properties for three such elements (scandium, gallium, and germanium). By 1886 all of these elements had been discovered and found to have properties very similar to those he had predicted.

Metals and Nonmetals

The diagonal line or stairway that starts to the left of boron in the periodic table (Figure 2.8) separates metals from nonmetals. The more than 80 elements to the left and below that line, shaded in blue in Figure 2.8, have the properties of **metals** (Figure 2.10a); in particular, they have high electrical conductivities. Elements above and to the right of the stairway shaded in yellow in Figure 2.8 are **nonmetals** (Figure 2.10b); about 18 elements fit in that category.

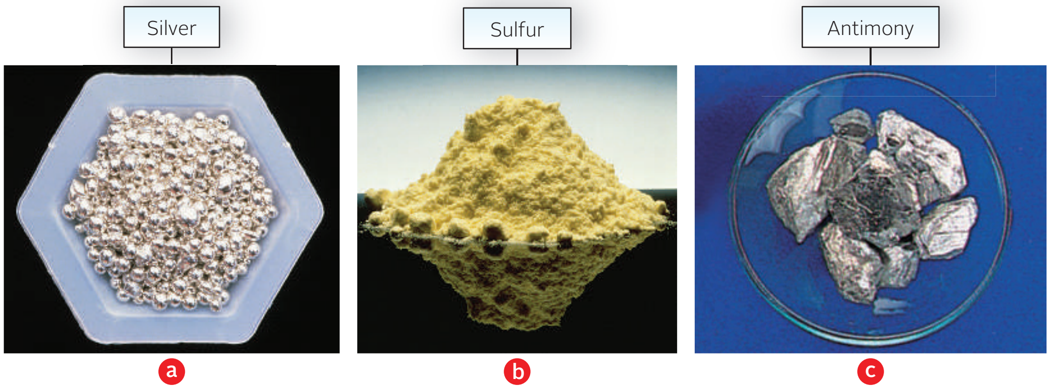
Along the stairway (zig-zag line) in the periodic table are several elements that are difficult to classify exclusively as metals or nonmetals. They have properties between those of elements in the two classes. In particular, their electrical conductivities are intermediate between those of metals and nonmetals. *i* The six elements shaded in green in Figure 2.8

i These elements, particularly Si, are used in semiconductors.

B	Si	Ge	As	Sb	Te
boron	silicon	germanium	arsenic	antimony	tellurium

are often called **metalloids** (Figure 2.10c).

Figure 2.10 Three elements. Silver (Group 11) is a metal. Sulfur (Group 16) is a nonmetal; antimony (Group 15) is a metalloid.



© Cengage Learning/Charles D. Winters (a-c)

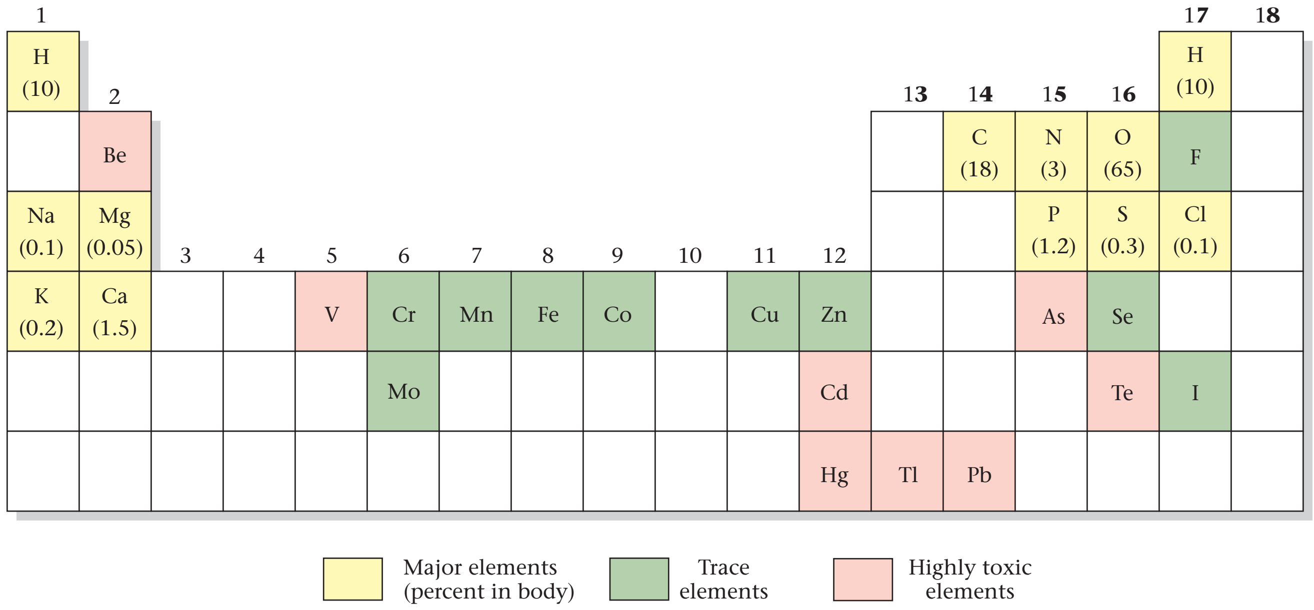


Figure 2.11 Biologically important elements and highly toxic elements. For the major elements in the body, their percent abundance in the body is given below the symbols.

Figure 2.11 shows the biologically important elements. The "good guys," essential to life, include the major elements (yellow), which account for 99.9% of total body mass, and the trace elements (green) required in very small quantities. In general, the abundances of elements in the body parallel those in the world around us, but there are some important exceptions. Aluminum and silicon, although widespread in nature, are missing in the human body. The reverse is true of carbon, which makes up only 0.08% of the earth's crust but 18% of the body, where it occurs in a variety of organic compounds including proteins, carbohydrates, and fats.

The "bad guys," shown in pink in Figure 2.11, are toxic, often lethal, even in relatively small quantities. Several of the essential trace elements *become* toxic if their concentrations in the body increase. Selenium is a case in point. You need about 0.00005 g/day to maintain good health, but 0.001 g/day can be deadly. That's a good thing to keep in mind if you're taking selenium supplements.

2-5 Molecules and Ions

Isolated atoms rarely occur in nature; only the noble gases (He, Ne, Ar, . . .) consist of individual, nonreactive atoms. Atoms tend to combine with one another in various ways to form more complex structural units. Two such units, which serve as building blocks for a great many elements and compounds, are molecules and ions.

Molecules

Two or more atoms may combine with one another to form an uncharged **molecule.** The atoms involved are usually those of nonmetallic elements. Within the molecule, atoms are held to one another by strong forces called *covalent bonds,* which consist of shared pairs of electrons (Chapter 7). Forces between neighboring molecules, in contrast, are quite weak.

Super Glue™ is weak compared with covalent bonds.

A **molecular substance** is most often represented by a **molecular formula,** in which the number of atoms of each element is indicated by a subscript written after the symbol of the element. Thus we interpret the molecular formulas for water (H_2O), ammonia (NH_3), and methane (CH_4) to mean that in

- the water molecule, there is one oxygen atom and two hydrogen atoms.
- the ammonia molecule, there is one nitrogen atom and three hydrogen atoms.
- the methane molecule, there is one carbon atom and four hydrogen atoms.

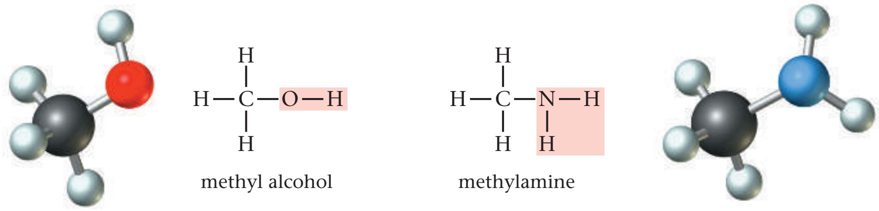
Structural formulas help us predict chemical properties.

The structure of a molecule is sometimes represented by a **structural formula,** which shows the bonding pattern within the molecule. The structural formulas of water, ammonia, and methane are

$$\mathrm{H{-}O{-}H} \qquad \begin{array}{c}\mathrm{H{-}N{-}H}\\ |\\ \mathrm{H}\end{array} \qquad \begin{array}{c}\mathrm{H}\\ |\\ \mathrm{H{-}C{-}H}\\ |\\ \mathrm{H}\end{array}$$

The dashes represent covalent bonds. The geometries of these molecules are shown in Figure 2.12.

Sometimes we represent a molecular substance with a formula intermediate between a structural formula and a molecular formula. A **condensed structural formula** suggests the bonding pattern in the molecule and highlights the presence of a reactive group of atoms within the molecule. Consider, for example, the organic compounds commonly known as methyl alcohol and methylamine. Their structural formulas are



The condensed structural formulas of these compounds are written as

$$CH_3\,OH \qquad CH_3NH_2$$

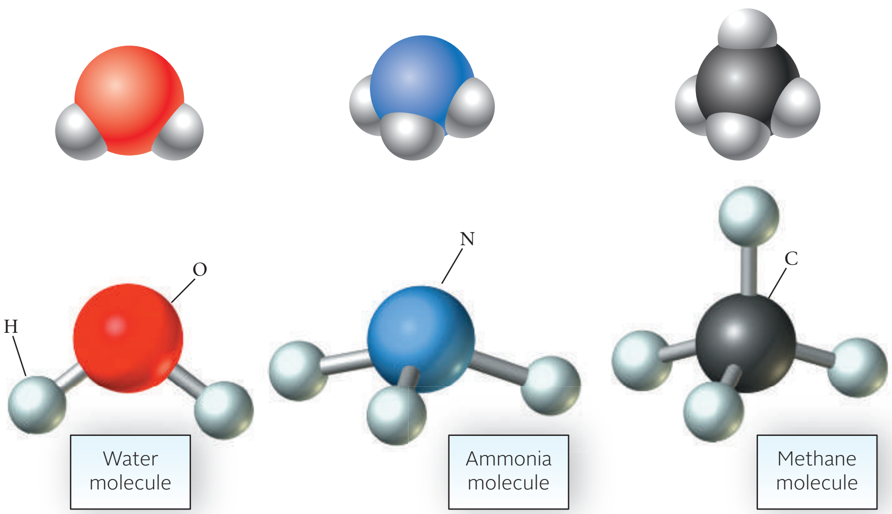


Figure 2.12 Space filling (top) and ball-and-stick (bottom) models of water (H_2O), ammonia (NH_3), and methane (CH_4). The "sticks" represent covalent bonds between H atoms and O, N, or C atoms. The models illustrate the geometry of the molecules (discussed in Chapter 7).

These formulas take up considerably less space and emphasize the presence in the molecule of

- the OH group found in all *alcohols,* including ethyl alcohol, CH_3CH_2OH, the alcohol found in intoxicating beverages such as beer and wine.
- the NH_2 group found in certain *amines,* including ethylamine, $CH_3CH_2NH_2$.

EXAMPLE 2.4

Give the molecular formulas of ethyl alcohol, CH_3CH_2OH, and ethylamine, $CH_3CH_2NH_2$.

ANALYSIS

Information given:	structural formula
Asked for:	molecular formula

STRATEGY

Add up the atoms of each element and use the sums as the subscripts for the element.

SOLUTION

ethyl alcohol	C: 1 + 1 = 2; H: 3 + 2 + 1 = 6; O: 1 molecular formula: C_2H_6O
ethylamine	C: 1 + 1 = 2; H: 3 + 2 + 2 = 7; N: 1 molecular formula: C_2H_7N

END POINT

Note that although molecular formulas give the composition of the molecule, they reveal nothing about the way the atoms fit together. In that sense they are less useful than the structural formulas.

Elements as well as compounds can exist as discrete molecules. In hydrogen gas, the basic building block is a molecule consisting of two hydrogen atoms joined by a covalent bond:

$$H—H$$

Other molecular elements are shown in Figure 2.13.

Ions

When an atom loses or gains electrons, charged particles called **ions** are formed. Metal atoms typically tend to lose electrons to form positively charged ions called **cations** (pronounced cát-ahy-uhn). Examples include the Na^+ and Ca^{2+} ions, formed from atoms of the metals sodium and calcium:

$$\underset{(11p^+,\ 11e^-)}{\text{Na atom}} \longrightarrow \underset{(11p^+,\ 10e^-)}{\text{Na}^+ \text{ ion}} + e^-$$

$$\underset{(20p^+,\ 20e^-)}{\text{Ca atom}} \longrightarrow \underset{(20p^+,\ 18e^-)}{\text{Ca}^{2+} \text{ ion}} + 2e^-$$

(The arrows separate *reactants,* Na and Ca atoms, from *products,* cations and electrons.)

Nonmetal atoms form negative ions called **anions** (pronounced án-ahy-uhn) by gaining electrons. Consider, for example, what happens when atoms of the nonmetals chlorine and oxygen acquire electrons:

$$\underset{(17p^+,\ 17e^-)}{\text{Cl atom}} + e^- \longrightarrow \underset{(17p^+,\ 18e^-)}{\text{Cl}^- \text{ ion}}$$

Group 15	Group 16	Group 17
		H_2 (*g*)
N_2 (*g*)	O_2 (*g*)	F_2 (*g*)
P_4 (*s*)	S_8 (*s*)	Cl_2 (*g*)
		Br_2 (*l*)
		I_2 (*s*)

Figure 2.13 Molecular elements and their physical states at room temperature: gaseous (*g*), liquid (*l*), or solid (*s*).

$$\text{O atom} + 2e^- \longrightarrow \text{O}^{2-} \text{ ion}$$
$$(8p^+, 8e^-) \qquad (8p^+, 10e^-)$$

C, P, and the metalloids do not form monatomic ions.

Notice that *when an ion is formed, the number of protons in the nucleus is unchanged.* It is the number of electrons that increases or decreases.

EXAMPLE 2.5

a Scandium, a transition metal, is added to aluminum to create an alloy used in sports equipment, like aluminum bats and bicycle frames. How many protons, neutrons, and electrons are there in this scandium ion: $^{45}_{21}Sc^{3+}$?

ANALYSIS

Information given:	nuclear symbol and charge
Information implied:	A, Z
Asked for:	p^+, n, e^-

STRATEGY (for all parts)

1. Recall the placement of Z and A in the nuclear symbol.
2. $Z = p^+$; $A = p^+ + n$; $e^- = p^+ -$ charge

SOLUTION

p^+, e^-	$^{45}_{21}Sc^{3+}$: $Z = p^+ = 21$; $e^- = p^+ - \text{(charge)} = 21 - (+3) = 18$;
n	$A = n + p^+$; $45 = n + 21$; $n = 24$

b Sulfur is present in an ore called chalcocite. The ion in the ore has 16 neutrons and 18 electrons. Write the nuclear symbol for the ion.

ANALYSIS

Information given:	e^-, n
Information implied:	Z
Asked for:	nuclear symbol for S

SOLUTION

Z,	From the periodic table, sulfur has atomic number 16 so $Z = 16 = p^+$.
A and charge	$A = p^+ + n = 16 + 16 = 32$; Charge $= p^+ - e^- = 16 - 18 = -2$
Nuclear symbol	$^{32}_{16}S^{2-}$

c A rare earth metal has varied uses from super computers to cancer therapy. It has an ion with a +2 charge. It has 37 electrons and 51 neutrons. Write its nuclear symbol.

ANALYSIS

Information given:	e^-, n, charge
Asked for:	nuclear symbol

SOLUTION

Z, element's identity	$Z = p^+ = e^- + \text{charge} = 37 + 2 = 39$; The element is Y(yttrium).
A	$A = p^+ + n = 39 + 51 = 90$ nuclear symbol: $^{A}_{Z}X = ^{90}_{39}Y^{2+}$

END POINT

A cation always contains more protons than electrons; the reverse is true of an anion.
Figure 2.14 shows the species described in this example.

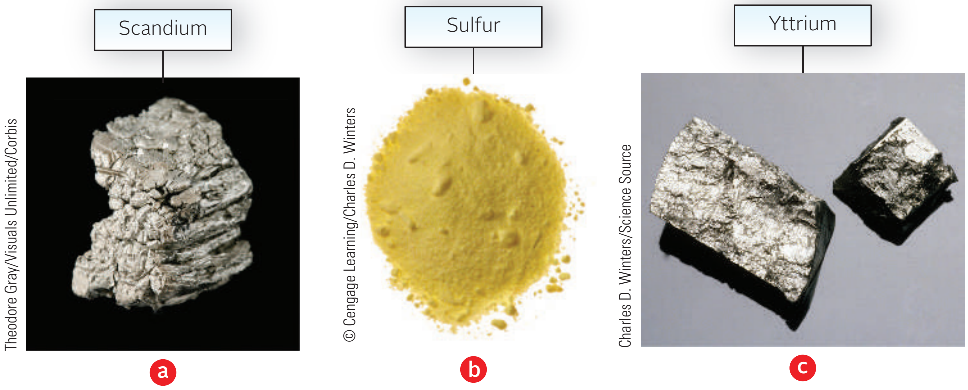


Figure 2.14 The ions of the elements shown are (a) Sc^{3+}, (b) S^{2-}, (c) Y^{2+}.

The ions dealt with to this point (e.g., Na^+, Cl^-) are *monatomic;* that is, they are derived from a single atom by the loss or gain of electrons. Many of the most important ions in chemistry are *polyatomic,* containing more than one atom. Examples include the hydroxide ion (OH^-) and the ammonium ion (NH_4^+). In these and other polyatomic ions, the atoms are held together by covalent bonds, for example,

We write +2 when describing the charge but 2+ when using it as a superscript in the formula of an ion.

$$(\mathrm{O{-}H})^- \qquad \left(\begin{array}{c} \mathrm{H} \\ | \\ \mathrm{H{-}N{-}H} \\ | \\ \mathrm{H} \end{array}\right)^+$$

In a very real sense, you can think of a polyatomic ion as a "charged molecule."

Because a bulk sample of matter is electrically neutral, any **ionic compound** always contains both cations (positively charged particles) and anions (negatively charged particles). Ordinary table salt, sodium chloride, is made up of an equal number of Na^+ and Cl^- ions. The structure of sodium chloride is shown in Figure 2.15. Notice that

- there are two kinds of structural units in NaCl, the Na^+ and Cl^- ions.
- there are no discrete molecules; Na^+ and Cl^- ions are bonded together in a continuous network.

Ionic compounds are held together by strong electrical forces between oppositely charged ions (e.g., Na^+, Cl^-). These forces are referred to as **ionic bonds**.

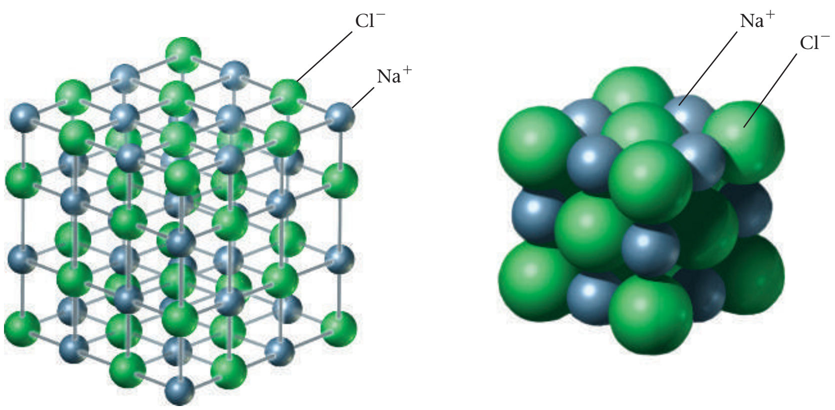


Figure 2.15 Sodium chloride structure. In these two ways of showing the structure, the small spheres represent Na^+ ions and the large spheres Cl^- ions. Note that in any sample of sodium chloride there are equal numbers of Na^+ and Cl^- ions, but no NaCl molecules.

Typically, ionic compounds are solids at room temperature and have relatively high melting points (mp NaCl = 801°C, $CaCl_2$ = 772°C). To melt an ionic compound requires that oppositely charged ions be separated from one another, thereby breaking ionic bonds.

When an ionic solid such as NaCl dissolves in water, the solution formed contains Na^+ and Cl^- ions. Since ions are charged particles, the solution conducts an electric current (Figure 2.16) and we say that NaCl is a strong *electrolyte*. In contrast, a water solution of sugar, which is a molecular solid, does not conduct electricity. Sugar and other molecular solutes are *nonelectrolytes*.

Then there are weak electrolytes, which dissolve mostly as molecules with a few ions.

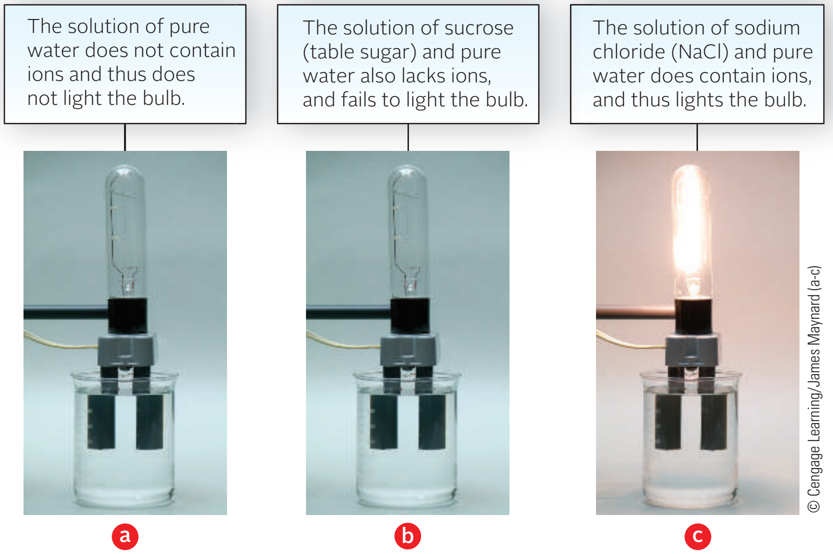
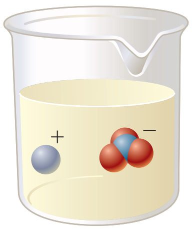


Figure 2.16 Electrical conductivity test. For electrical current to flow and light the bulb, the solution in which the electrodes are immersed must contain ions, which carry electrical charge.

EXAMPLE 2.6 CONCEPTUAL

The structure of a water solution of KNO_3, containing equal numbers of K^+ and NO_3^- ions, may be represented as

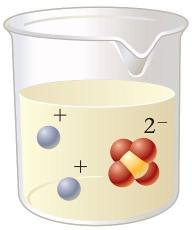


(H_2O molecules are not shown.) Construct a similar beaker to show the structure of a water solution of potassium sulfate, K_2SO_4. Use a yellow sphere to represent a sulfur atom.

STRATEGY

Once you deduce the formula of the ionic compound, it's all downhill. Remember, though, that you have to show the relative numbers of cations and anions.

SOLUTION



2-6 Formulas of Ionic Compounds

When a metal such as sodium (Na) or calcium (Ca) reacts with a nonmetal such as chlorine (Cl_2), the product is ordinarily an ionic compound. The formula of that compound (e.g., NaCl, $CaCl_2$) shows the simplest ratio between cation and anion (one Na^+ ion for one Cl^- ion; one Ca^{2+} ion for two Cl^- ions). In that sense, the formulas of ionic compounds are simplest formulas. Notice that the symbol of the metal (Na, Ca) always appears first in the formula, followed by that of the nonmetal.

To predict the formula of an ionic compound, you need to know the charges of the two ions involved. Then you can apply the principle of **electrical neutrality**, which requires that ***the total positive charge of the cations in the formula must equal the total negative charge of the anions.*** Consider, for example, the ionic compound calcium chloride. The ions present are Ca^{2+} and Cl^-. For the compound to be electrically neutral, there must be two Cl^- ions for every Ca^{2+} ion. The formula of calcium chloride must be $CaCl_2$, indicating that the simplest ratio of Cl^- to Ca^{2+} ions is 2:1.

Cations and Anions with Noble-Gas Structures

The charges of ions formed by atoms of the main-group elements can be predicted by applying a simple principle:

> ***Atoms that are close to a noble gas*** (Group 18) ***in the periodic table form ions that contain the same number of electrons as the neighboring noble-gas atom.***

This is reasonable; noble-gas atoms must have an extremely stable electronic structure, because they are so unreactive. Other atoms might be expected to acquire noble-gas electronic structures by losing or gaining electrons.

Applying this principle, you can deduce the charges of ions formed by main-group atoms:

Group	No. of Electrons in Atom	Charge of Ion Formed	Examples
1	1 more than noble-gas atom	+1	Na^+, K^+
2	2 more than noble-gas atom	+2	Mg^{2+}, Ca^{2+}
16	2 less than noble-gas atom	−2	O^{2-}, S^{2-}
17	1 less than noble-gas atom	−1	F^-, Cl^-

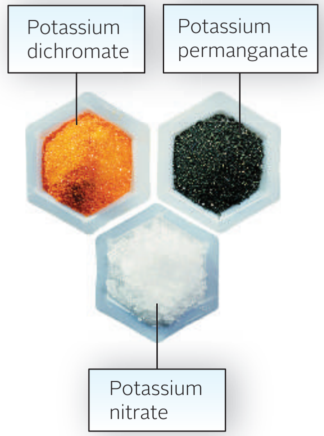
Two other ions that have noble-gas structures are

Al^{3+} (Al has three more e^- than the preceding noble gas, Ne)

N^{3-} (N has three fewer e^- than the following noble gas, Ne)

Table 2.2 Some Common Polyatomic Ions

+1	−1	−2	−3
NH_4^+ (ammonium)	OH^- (hydroxide)	CO_3^{2-} (carbonate)	PO_4^{3-} (phosphate)
Hg_2^{2+} (mercury I)	NO_3^- (nitrate)	SO_4^{2-} (sulfate)	
	ClO_3^- (chlorate)	CrO_4^{2-} (chromate)	
	ClO_4^- (perchlorate)	$Cr_2O_7^{2-}$ (dichromate)	
	CN^- (cyanide)	HPO_4^{2-} (hydrogen phosphate)	
	$C_2H_3O_2^-$ (acetate)		
	MnO_4^- (permanganate)		
	HCO_3^- (hydrogen carbonate)		
	$H_2PO_4^-$ (dihydrogen phosphate)		



© Cengage Learning/Charles D. Winters

Figure 2.17 Ionic compounds containing polyatomic ions. Potassium dichromate ($K_2Cr_2O_7$), potassium permanganate ($KMnO_4$), and potassium nitrate (KNO_3).

Cations of the Transition and Post-Transition Metals

Several metals that are farther removed from the noble gases in the periodic table form positive ions. These include the transition metals in Groups 3 to 12 and the post-transition metals in Groups 13 to 15. The cations formed by these metals typically have charges of +1, +2, or +3 and ordinarily do not have noble-gas structures. We will postpone to Chapter 4 a general discussion of the specific charges of cations formed by these metals.

Many of the transition and post-transition metals form more than one cation. Consider, for example, iron in Group 8. This metal forms two different series of compounds with nonmetals. In one series, iron is present as a +2 cation

$$Fe^{2+}: \quad FeCl_2, FeBr_2, \ldots$$

In the other series, iron exists as a +3 cation

$$Fe^{3+}: \quad FeCl_3, FeBr_3, \ldots$$

Polyatomic Ions

Table 2.2 lists some of the polyatomic ions that you will need to know, along with their names and charges. Figure 2.17 shows some of the compounds that have polyatomic ions. Notice that

- there are only two common polyatomic cations, NH_4^+ and Hg_2^{2+}. ***All other cations considered in this text are derived from individual metal atoms*** (e.g., Na^+ from Na, Ca^{2+} from Ca, . . .).
- most of the polyatomic anions contain one or more oxygen atoms; collectively these species are called **oxoanions**.

EXAMPLE 2.7

Predict the formula of the ionic compound

(a) formed by barium with iodine.

(b) containing a transition metal with a +1 charge in period 4 and Group 11 and oxide ions.

(c) containing an alkaline earth in period 5 and nitrogen.

(d) containing ammonium and phosphate ions.

STRATEGY

1. Recall charge of metals: group 1 (+1); group 2 (+2); Al (+3)
2. Recall charge of nonmetals: group 16: −2; group 17: −1; N: −3
3. The formula has to be electrically neutral.
4. Use Table 2.2 for polyatomic ions.

SOLUTION

(a) Charges	Ba is in group 2; thus its charge is +2. I is in group 17, thus its charge is −1.
electrical neutrality	$Ba^{2+}I^{1-}$; $2I^-$ are needed. The formula is $\boxed{BaI_2}$.
(b) Charges	Period 4, group 4 is Cu with a given charge of +1. O is in group 16, thus its charge is −2.
electrical neutrality	$Cu^{1+}O^{2-}$; 2 Cu^{1+} are needed. The formula is $\boxed{Cu_2O}$.
(c) Charges	The alkaline earth (group 2) in period 5 is Sr; thus its charge is +2. N in an ionic compound is always −3.
electrical neutrality	$Sr^{2+}N^{3-}$; 3 Sr^{2+} and 2 N^{3-} are needed. The formula is $\boxed{Sr_3N_2}$.
(d) Charges	Ammonium is a polyatomic ion: NH_4^+ (Table 2.2). Phosphate is a polyatomic ion: PO_4^{3-} (Table 2.2).
electrical neutrality	NH_4^+ PO_4^{3-}; 3 NH_4^+ are needed. The formula is $\boxed{(NH_4)_3PO_4}$.

continued

END POINT

To be able to write the formulas of compounds, you must know the symbols of the elements. You must also know the symbols and charges of the polyatomic ions listed in Table 2.2. Learn them soon!

2-7 Names of Compounds

A compound can be identified either by its formula (e.g., NaCl) or by its name (sodium chloride). In this section, you will learn the rules used to name ionic and simple molecular compounds. To start with, it will be helpful to show how individual ions within ionic compounds are named.

i To name a compound, you have to decide whether it is molecular or ionic; the rules are different.

Ions

Monatomic cations take the name of the metal from which they are derived. Examples include

Na^+ sodium K^+ potassium

There is one complication. As mentioned earlier, certain metals in the transition and post-transition series form more than one cation, for example, Fe^{2+} and Fe^{3+}. To distinguish between these cations, the charge must be indicated in the name. This is done by putting the charge as a Roman numeral in parentheses after the name of the metal:

Fe^{2+} iron(II) Fe^{3+} iron(III)

(An older system used the suffixes *-ic* for the ion of higher charge and *-ous* for the ion of lower charge. These were added to the stem of the Latin name of the metal, so that the Fe^{3+} ion was referred to as ferric and the Fe^{2+} ion as ferrous.)

Monatomic anions are named by adding the suffix *-ide* to the stem of the name of the nonmetal from which they are derived.

				H^-	hydride
N^{3-}	nitride	O^{2-}	oxide	F^-	fluoride
		S^{2-}	sulfide	Cl^-	chloride
		Se^{2-}	selenide	Br^-	bromide
		Te^{2-}	telluride	I^-	iodide

Polyatomic ions, as you have seen (Table 2.2), are given special names. Certain nonmetals in Groups **15** to **17** of the periodic table form more than one polyatomic ion containing oxygen (oxoanions). The names of several such oxoanions are shown in Table 2.3. From the entries in the table, you should be able to deduce the following rules:

1. When a nonmetal forms two oxoanions, the suffix *-ate* is used for the anion with the larger number of oxygen atoms. The suffix *-ite* is used for the anion containing fewer oxygen atoms.
2. When a nonmetal forms more than two oxoanions, the prefixes *per-* (largest number of oxygen atoms) and *hypo-* (fewest oxygen atoms) are used as well.

Ionic Compounds

The name of an ionic compound consists of two words. The first word names the cation and the second names the anion. This is, of course, the same order in which the ions appear in the formula.

Cl, Br, and I form more than two oxoanions.

The oxoanions of bromine and iodine are named like those of chlorine.

Table 2.3 Oxoanions of Nitrogen, Sulfur, and Chlorine

Nitrogen	Sulfur	Chlorine
		ClO_4^- *per*chlor*ate*
NO_3^- nitr*ate*	SO_4^{2-} sulf*ate*	ClO_3^- chlor*ate*
NO_2^- nitr*ite*	SO_3^{2-} sulf*ite*	ClO_2^- chlor*ite*
		ClO^- *hypo*chlor*ite*

In naming the compounds of transition or post-transition metals, we ordinarily indicate the charge of the metal cation by a Roman numeral:

$Cr(NO_3)_3$ chromium(III) nitrate $SnCl_2$ tin(II) chloride

In contrast, we never use Roman numerals with compounds of the Group 1 or Group 2 metals; they always form cations with charges of +1 or +2, respectively.

EXAMPLE 2.8

Name the following ionic compounds:

(a) CaS (b) $Al(NO_3)_3$ (c) $FeCl_2$

STRATEGY

Recall symbols for elements, symbols for polyatomic ions (Table 2.2), and suffixes for nonmetals.

SOLUTION

(a) CaS	Ca = calcium; S = sulfur → sulfide; calcium sulfide
(b) $Al(NO_3)_3$	Al^{3+} = aluminum; NO_3^- = nitrate; aluminum nitrate
(c) $FeCl_2$	Fe^{2+} = iron, which is a transition metal so (II) should be written after the name of the metal; Cl^- = chlorine → chloride; iron(II) chloride

© Cengage Learning/Charles D. Winters

Nitrogen dioxide (NO_2), a binary molecular compound. It is a reddish-brown gas at 25°C and 1 atm.

Binary Molecular Compounds

When two nonmetals combine with each other, the product is most often a binary molecular compound. There is no simple way to deduce the formulas of such compounds. There is, however, a systematic way of naming molecular compounds that differs considerably from that used with ionic compounds.

The systematic name of a binary molecular compound, which contains two different nonmetals, consists of two words.

1. The first word gives the name of the element that appears first in the formula; a Greek prefix (Table 2.4) is used to show the number of atoms of that element in the formula.
2. The second word consists of
   - the appropriate Greek prefix designating the number of atoms of the second element. (See Table 2.4.)
   - the stem of the name of the second element.
   - the suffix *-ide*.

Table 2.4 Greek Prefixes Used in Nomenclature

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

*The prefix mono (1) is seldom used.

To illustrate these rules, consider the names of the several oxides of nitrogen:

N_2O_5	*di*nitrogen *penta*ox*ide*	N_2O_3	*di*nitrogen *tri*ox*ide*
N_2O_4	*di*nitrogen *tetra*ox*ide*	NO	nitrogen ox*ide*
NO_2	nitrogen *di*ox*ide*	N_2O	*di*nitrogen ox*ide*

When the prefixes tetr*a*, pent*a*, hex*a*, . . . are followed by the letter "o," the *a* is often dropped. For example, N_2O_5 is often referred to as *dinitrogen pentoxide*.

Many of the best-known binary compounds of the nonmetals have acquired common names. These are widely—and in some cases exclusively—used. Examples include

H_2O	water	PH_3	phosphine
H_2O_2	hydrogen peroxide	AsH_3	arsine
NH_3	ammonia	NO	nitric oxide
N_2H_4	hydrazine	N_2O	nitrous oxide
C_2H_2	acetylene	CH_4	methane

EXAMPLE 2.9

Give the names of the following molecules:

(a) SF_4 (b) PCl_3 (c) N_2O_3 (d) Cl_2O_7

STRATEGY

1st element: subscript → prefix (Table 2.4) + element name

2nd element: subscript → prefix + element name ending in *ide*

SOLUTION

(a) SF_4	S: subscript = 1; no prefix = sulfur F: subscript = 4 = tetra; F = fluorine → fluoride; sulfur tetrafluoride
(b) PCl_3	P: subscript = 1; no prefix = phosphorus Cl: subscript = 3 = tri; Cl = chlorine → chloride; phosphorus trichloride
(c) N_2O_3	N: subscript = 2 = di; dinitrogen O: subscript = 3 = tri; O = oxygen → oxide; dinitrogen trioxide
(d) Cl_2O_7	Cl: subscript = 2 = di; dichlorine O: subscript = 7 = hepta; O = oxygen → oxide; dichlorine heptaoxide

Acids

A few binary molecular compounds containing H atoms ionize in water to form H^+ ions. These are called *acids*. One such compound is hydrogen chloride, HCl; in water solution it exists as aqueous H^+ and Cl^- ions. The water solution of hydrogen chloride is given a special name: It is referred to as *hydrochloric acid*. A similar situation applies with HBr and HI:

Pure Substance		Water Solution	
$HCl(g)$	Hydrogen chloride	$H^+(aq)$, $Cl^-(aq)$	Hydrochloric acid
$HBr(g)$	Hydrogen bromide	$H^+(aq)$, $Br^-(aq)$	Hydrobromic acid
$HI(g)$	Hydrogen iodide	$H^+(aq)$, $I^-(aq)$	Hydriodic acid

Most acids contain oxygen in addition to hydrogen atoms. Such species are referred to as **oxoacids.** Two oxoacids that you are likely to encounter in the general chemistry laboratory are

HNO_3 nitric acid H_2SO_4 sulfuric acid

The names of oxoacids are simply related to those of the corresponding oxoanions. The *-ate* suffix of the anion is replaced by *-ic* in the acid. In a similar way, the suffix *-ite* is replaced by the suffix *-ous*. The prefixes *per-* and *hypo-* found in the name of the anion are retained in the name of the acid.

ClO_4^-	*per*chlor*ate* ion	$HClO_4$	*per*chlor*ic* acid
ClO_3^-	chlor*ate* ion	$HClO_3$	chlor*ic* acid
ClO_2^-	chlor*ite* ion	$HClO_2$	chlor*ous* acid
ClO^-	*hypo*chlor*ite* ion	$HClO$	*hypo*chlor*ous* acid

Figures 2.18 and 2.19 summarize the rules for naming ionic compounds, molecules, and acids.

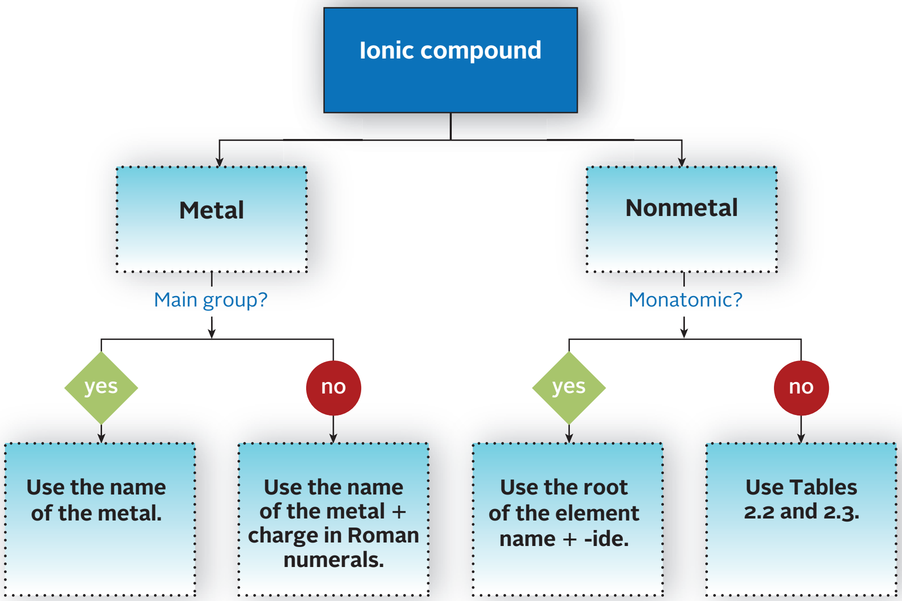


Figure 2.18 Flowchart for naming binary ionic compounds.

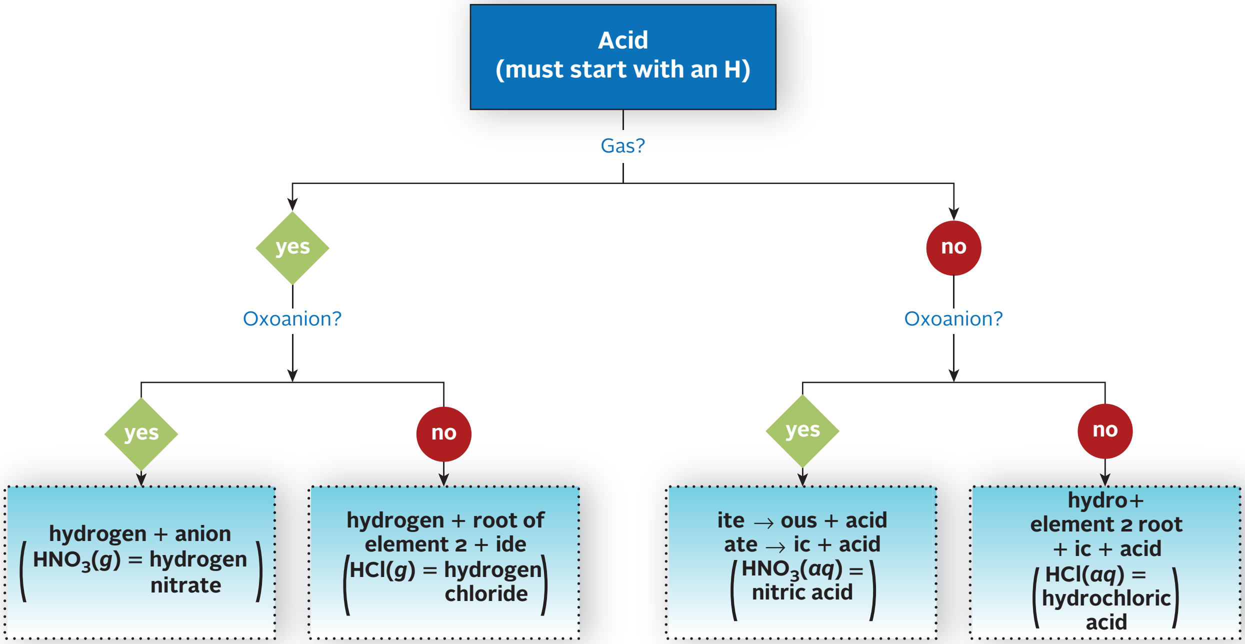


Figure 2.19 Flow chart for determining an acid.

EXAMPLE 2.10

Give the names of

a $HBr(g)$

STRATEGY

Gases follow the rules for naming binary molecules.

SOLUTION

$HBr(g)$	No prefixes for both elements (subscripts are both 1). H = hydrogen; Br = bromine → bromide; **hydrogen bromide**

b $ClO_2(aq)$

STRATEGY

Name of oxoanion (Table 2.3) (change *ite* to *ous*) + the word "acid."

SOLUTION

$ClO_2(aq)$	ClO_2^- = chlorite → chlorous + acid; **chlorous acid**

c $H_2SO_4(aq)$

STRATEGY

Name of oxoanion (Table 2.3) (change *ate* to *ic*) + the word "*acid*."

SOLUTION

$H_2SO_4(aq)$	SO_4^{2-} = sulfate → sulfuric + acid; **sulfuric acid**

continued

d HIO(aq)

STRATEGY

Name of oxoanion (Table 2.3) (change *ite* to *ous*) + the word "acid." The oxoanion's name is analogous to the naming of the chlorine oxoanions.

SOLUTION

HIO(aq)	IO^- = hypoiodite → hypoiodous + acid; hypoiodous acid Hypoiodous is analogous to hypochlorous, the name for ClO^-.

END POINT

You need to learn the names of the oxoanions listed in Table 2.3.

CHEMISTRY BEYOND THE CLASSROOM

Mastering the Peri'god'ic Table

by Rob Kyff

In ancient Rome, Mercury (the messenger of the gods) was constantly zipping around from deity to deity. So when the Romans needed a word for the poisonous metallic element that flowed quickly at room temperature, they named it "Mercury" for their speedy courier (Figure A).

In fact, when it comes to naming elements, Greek and Roman gods pretty much ran the (periodic) table.

When a new planet swam into the ken of British astronomer William Herschel in 1781, for instance, he named it "Uranus," after the Roman god of the sky. Seven years later, when the German chemist Martin Klaproth discovered a radioactive new element, he decided it would be trendy and hip to name it after the new planet—voila!—"uranium" (Figure B).

And when English clergyman William Gregor detected another new metallic element in 1791, Klaproth was again on the case. He wanted to name the newcomer for a child of Uranus, but somehow "Uranus, Jr." just didn't work.

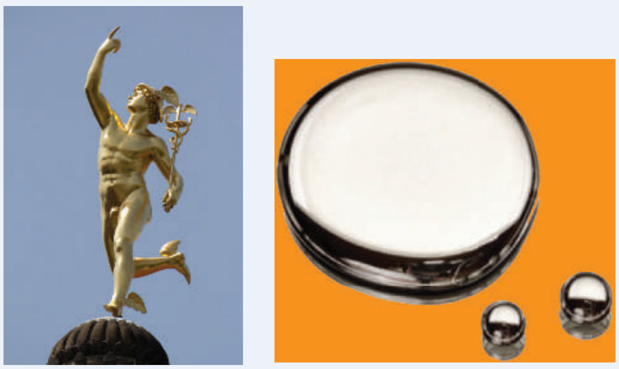
So he named it for all of Uranus' children: a race of giants known as "the Titans." And so "titanium" was born.

Seven years later, obviously on a roll, Klaproth dubbed another new element "tellurium," for Tellus, a.k.a. "Terra," the Roman goddess of earth.

In 1801, Sicilian astronomer Giuseppi Piazzi spotted what seemed to be a strange, small planet orbiting the sun between Mars and Jupiter. Because Ceres, the Roman goddess of agriculture, had been viewed as the protector of Piazzi's native Sicily, he named the new object, which we now know to be an asteroid, "Ceres."

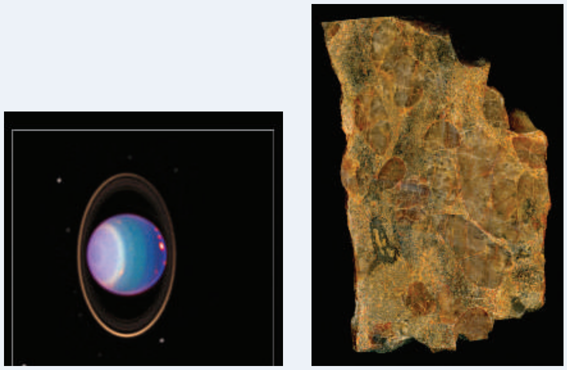
So just two years after Piazzi's discovery, the existence of another new element was verified. You can guess what hap-

Source: By permission of Rob Kyff and Creators Syndicate, Inc



Left, © Rob Wilson/Shutterstock.com; *right*, © Cengage Learning/Charles D. Winters.com

Figure A (*Left*) Statue of the Roman god Mercury by the German sculptor Ludwig von Hofer, on the top of the Old Chancellery Building in Stuttgart, Germany. (*Right*) The element mercury.



Left, NASA/JPL/STScI; *right*, TED KINSMAN/SCIENCE PHOTO LIBRARY

Figure B (*Left*) A false-color image of the planet Uranus, as seen from the Hubble Space Telescope. (*Right*) Uranium ore conglomerate (also called pitchblend). The dark areas between the large quartz pebbles are uranium.

pened next. Klaproth, always eager to mimic the god-given nomenclature of astronomers, called the new element "Cerium."

Then in 1868, with Klaproth gone but his naming methods not forgotten, British astronomer Norman Lockyer noticed that one element detected in the sun's spectrum could not be correlated with any element found on Earth. So he named this new, solar-generated element "helium" for "helios," the Greek god of the sun.

When Tantalus, a mortal son of Zeus, shared his father's secrets with ordinary humans, his father punished him by placing him in fresh water up to his chin and dangling luscious fruits above his head. Doesn't sound so bad, right? But every time Tantalus stooped to quaff his thirst, the water drained away. And when he reached for the fruit, it moved upward beyond his grasp. This excruciating torture inspired the English word "tantalize."

In 1814, the Swedish chemist Anders Eckeberg discovered a new metal that could be placed in strong acid without "drinking" it, that is, reacting with it. So his fellow Swedish chemist Jöns Berzelius, remembering the perpetually parched Tantalus, dubbed the new element "tantalum."

At about the same time, the English chemist Charles Hatchett discovered a new metal, in Connecticut. He dubbed it "columbium," after "Columbia," a term for the United States. But other scientists claimed that columbium was identical to tantalum and wasn't new at all.

This dispute was resolved in 1846 when German chemist Heinrich Rose determined that columbium was indeed a new element. But because it resembled tantalum so closely, Rose dubbed it "niobium" after Tantalus' weeping daughter, Niobe. American scientists stuck with "columbium" for a while, but "niobium" eventually prevailed.

Another mythological daughter also gave her name to an element. Selene, the daughter of Hyperion, was the Greek goddess of the moon.

In 1818 the aforementioned Berzelius discovered a new element very similar to tellurium. Because "tellurium" had been named for Tellus, the Roman goddess of the earth, Berzelius strove for balance by calling its sister element "selenium" for the goddess of the moon.

Finally, in 1940 chemists at the University of California produced two new radioactive elements derived from uranium (which had been named in 1788 for Uranus, the Roman god of the sky).

So, following the order of the planets, they named the newcomers "neptunium" (for Neptune, Roman god of the sea) and "plutonium" (for Pluto, Greek god of the underworld).

Though Pluto is no longer considered a planet, its namesake element abides.

Key Concepts

1. Relate a nuclear symbol to the number of protons and neutrons.
   (**Example 2.1; Problems 7–16**)
2. Relate atomic mass, isotopic abundance, and average mass of an element.
   (**Example 2.2; Problems 17–30**)
3. Relate atomic mass to Avogadro's number.
   (**Example 2.3; Problems 31–38**)
4. Relate elements and the periodic table.
   (**Problems 39–46**)
5. Relate structural, condensed, and molecular formulas.
   (**Example 2.4; Problems 47, 48**)
6. Relate the ionic charge to the number of electrons.
   (**Example 2.5; Problems 49–52**)
7. Predict formulas of ionic compounds from charge of ions.
   (**Example 2.7; Problems 59, 60**)
8. Relate names to formulas
   - Ionic compounds
   (**Example 2.8; Problems 61–64**)
   - Binary molecular compounds
   (**Example 2.9; Problems 55–57**)
   - Oxoacids and oxoanions
   (**Example 2.10; Problems 65, 66**)

Key Equation

$$\text{Atomic Mass Y} = (\text{atomic mass Y}_1) \times \frac{\%\ \text{Y}_1}{100\%} + (\text{atomic mass Y}_2) \times \frac{\%\ \text{Y}_2}{100\%} + \cdots$$

Key Terms

atom
atomic mass
atomic mass units (amu)
atomic number
Avogadro's number
electrons
formula
—condensed structural formula
—molecular formula
—structural formula
ions
—anions
—cations
isotope
main-group elements
mass number
metals
metalloids
molecule
neutron
nonmetals
nuclear symbol
nucleus
oxoacids
oxoanions
periodic table
—groups
—periods
proton
transition metals

Summary Problem

Sulfur is a bright yellow solid at room temperature. It is one of the components in gunpowder, matches, and fireworks. It is added to rubber to make it tougher (vulcanization) and to topical ointments for the treatment of skin disorders.

(a) What is its molecular formula?

(b) How many protons and electrons are there in an atom of sulfur?

(c) If aluminum ions combine with sulfide ions, is the resulting compound molecular? What is the compound's name and formula?

(d) Is the compound described in (c) an electrolyte?

(e) Two atoms of sulfur combine with two atoms of chlorine. Is this compound a molecule? What is its name and formula?

(f) Write the nuclear symbol for a sulfur atom with 16 protons and 18 neutrons.

(g) To what group and period in the periodic table does sulfur belong?

(h) How many neutrons are there in the isotope S-36?

(i) The naturally occurring isotopes of sulfur, their atomic masses, and abundances are:
S-32: 31.97207 amu - 94.93%
S-33: 32.97146 amu - 0.76%
S-34: 33.96787 amu - 4.29%
S-36: 35.96708 amu - 0.02%
What is the average atomic mass for sulfur?

(j) How many atoms are there in 12.55 g of elemental sulfur (S) ?

(k) What is the mass of a billion (1×10^9) atoms of sulfur?

(l) Sulfur can combine with oxygen in many different ways. Write the names of the following compounds and ions: SO_3, H_2SO_3 (*aq*), SO_4^{2-}, Na_2SO_3.

Answers

(a) S_8
(b) **16 protons, 16 electrons**
(c) **no; Al_2S_3 - aluminum sulfide**
(d) **yes**
(e) **yes; S_2Cl_2 - disulfur dichloride**
(f) $^{34}_{16}S$
(g) **group 16, period 3**
(h) **20 neutrons**
(i) **32.07 amu**
(j) **2.357×10^{23} atoms**
(k) **5.325×10^{-14} g**
(l) **SO_3: sulfur trioxide**
H_2SO_3: sulfurous acid
SO_4^{2-}: sufate ion
Na_2SO_3: sodium sulfite

Questions and Problems

Even-numbered questions and Challenge Problems have answers in Appendix 5 and fully worked solutions in the *Student Solutions Manual*.

2-1 Atoms and the Atomic Theory

Atomic Theory and Laws

1. State in your own words the law of conservation of mass. State the law in its modern form.

2. State in your own words the law of constant composition.

3. Two basic laws of chemistry are the law of conservation of mass and the law of constant composition. Which of these laws (if any) do the following statements illustrate?

(a) Lavoisier found that when mercury(II) oxide, HgO, decomposes, the total mass of mercury (Hg) and oxygen formed equals the mass of mercury(II) oxide decomposed.

(b) Analysis of the calcium carbonate found in the marble mined in Carrara, Italy, and in the stalactites of the Carlsbad Caverns in New Mexico gives the same value for the percentage of calcium in calcium carbonate.

(c) Hydrogen occurs as a mixture of two isotopes, one of which is twice as heavy as the other.

4. Two basic laws of chemistry are the law of conservation of mass and the law of constant composition. Which of these laws (if any) do the following statements illustrate?

(a) The mass of phosphorus, P, combined with one gram of hydrogen, H, in the highly toxic gas phosphene, PH_3, is a little more than twice the mass of nitrogen, N, combined with one gram of hydrogen in ammonia gas, NH_3.

(b) A cold pack has the same mass before and after the seal between two reactants is broken to allow reaction to occur.

(c) It is highly improbable that carbon monoxide gas found in Los Angeles is $C_{1.2}O_{1.1}$.

2-2 Components of the Atom

5. Who discovered the electron? Describe the experiment that led to the deduction that electrons are negatively charged particles.

6. Who discovered the nucleus? Describe the experiment that led to this discovery.

7. Selenium is widely sold as a dietary supplement. It is advertised to "protect" women from breast cancer. Write the nuclear symbol for naturally occurring selenium. It has 34 protons and 46 neutrons.

8. Radon is a radioactive gas that can cause lung cancer. It has been detected in the basement of some homes. In an atom of Rn-220,

(a) how many protons are there?

(b) how many neutrons are there?

2-3 Quantitative Properties of the Atom

Nuclear Symbols and Isotopes

9. How do the isotopes of argon, Ar-36, Ar-38, and Ar-40, differ from each other? Write the nuclear symbol for all three isotopes.

10. Consider two isotopes Fe-54 and Fe-56.

(a) Write the nuclear symbol for both isotopes.

(b) How do they differ from each other?

11. Uranium-235 is the isotope of uranium commonly used in nuclear power plants. How many

(a) protons are in its nucleus?

(b) neutrons are in its nucleus?

(c) electrons are in a uranium atom?

12. An isotope of americium (Am) with 146 neutrons is used in many smoke alarms.

(a) How many electrons does an atom of americium have?

(b) What is the isotope's mass number *A*?

(c) Write its nuclear symbol.

13. Consider the following nuclear symbols. How many protons, neutrons, and electrons does each element have? What elements do R, T, and X represent?

(a) ${}^{30}_{14}R$ (b) ${}^{89}_{39}T$ (c) ${}^{133}_{55}X$

14. Consider the following nuclear symbols. How many protons, neutrons, and electrons does each element have? What elements do A, L, and Z represent?

(a) ${}^{75}_{33}A$ (b) ${}^{51}_{23}L$ (c) ${}^{131}_{54}Z$

15. Nuclei with the same mass number but different atomic numbers are called *isobars*. Consider Ca-40, Ca-41, K-41 and Ar-41.

(a) Which of these are isobars? Which are isotopes?

(b) What do Ca-40 and Ca-41 have in common?

(c) Correct the statement (if it is incorrect): Atoms of Ca-41, K-41, and Ar-41 have the same number of neutrons.

16. See the definition for isobars in Question 15. Consider Cr-54, Fe-54, Fe-58, and Ni-58.

(a) Which of these are isobars? Which are isotopes?

(b) What do Fe-54 and Fe-58 have in common?

(c) Which atoms have the same number of neutrons?

Atomic Masses and Isotopic Abundances

17. Calculate the mass ratio of a bromine atom to an atom of

(a) neon (b) calcium (c) helium

18. Arrange the following in the order of increasing mass.

(a) a potassium ion, K^+

(b) a phosphorus molecule, P_4

(c) a potassium atom

(d) a platinum atom

19. Cerium is the most abundant rare earth metal. Pure cerium ignites when scratched by even a soft object. It has four known isotopes: ^{136}Ce (atomic mass = 135.907 amu), ^{138}Ce (atomic mass = 137.905 amu), ^{140}Ce (atomic mass = 139.905 amu), and ^{142}Ce (atomic mass = 141.909 amu). Ce-140 and Ce-142 are fairly abundant. Which is the more abundant isotope?

20. Consider the three stable isotopes of oxygen with their respective atomic masses: O-16 (15.9949 amu), O-17 (16.9993 amu), O-18 (17.9992 amu). Which is the most abundant?

21. Bromine has two occuring isotopes: ^{79}Br with atomic mass 78.9183 and ^{81}Br with atomic mass 80.9163. Without using a calculator, what would you estimate the % abundance of Br-79 to be?

22. Rubidium has two naturally occurring isotopes: ^{85}Rb (atomic mass = 84.9118 amu) and ^{87}Rb (atomic mass = 86.9092 amu). The percent abundance of ^{87}Rb can be estimated to be which of the following?

(a) 0% (b) 25% (c) 50% (d) 75%

23. Strontium has four isotopes with the following masses: 83.9134 amu (0.56%), 85.9094 amu (9.86%), 86.9089 amu (7.00%), and 87.9056 amu (82.58%). Calculate the average atomic mass of strontium.

24. Neon is an inert gas with three stable isotopes. It is used in gas lasers and in advertising signs. Its isotopes and their abundances are:

Ne-20	19.9924 amu	90.51%
Ne-21	20.9938 amu	0.27%
Ne-22	21.9914 amu	9.22%

What is the average atomic mass of neon?

25. Naturally occurring silver (Ag) consists of two isotopes. One of the isotopes has a mass of 106.90509 amu and 51.84% abundance. What is the atomic mass of the other isotope?

26. Copper has two naturally occurring isotopes. Cu-63 has an atomic mass of 62.9296 amu and an abundance of 69.17%. What is the atomic mass of the second isotope? What is its nuclear symbol?

27. Silicon (average atomic mass = 28.0855 amu) has three isotopes. Their masses are 27.9769 amu, 28.9765 amu, and 29.9738 amu. The abundance of the heaviest isotope is 2.96%. Estimate the abundances of the first two isotopes.

28. Magnesium (average atomic mass = 24.305 amu) consists of three isotopes with masses 23.9850 amu, 24.9858 amu, and 25.9826 amu. The abundance of the middle isotope is 10.00%. Estimate the abundances of the other isotopes.

29. Zinc has four stable isotopes: Zn-64, Zn-66, Zn-67, and Zn-68. Their natural abundances are 48.6%, 27.9%, 4.1%, and 18.8%, respectively. Sketch the mass spectrum for zinc.

30. Chlorine has two isotopes, Cl-35 and Cl-37. Their abundances are 75.53% and 24.47%, respectively. Assume that the only hydrogen isotope present is H-1.

(a) How many different HCl molecules are possible?
(b) What is the sum of the mass numbers of the two atoms in each molecule?
(c) Sketch the mass spectrum for HCl if all the positive ions are obtained by removing a single electron from an HCl molecule.

31. Lead is a heavy metal that remains in the bloodstream, causing mental retardation in children. It is believed that 3×10^{-7} g of Pb in 1.00 mL of blood is a health hazard. For this amount of lead how many atoms of lead are there in one mL of a child's blood?

32. Silversmiths are warned to limit their exposure to silver in the air to 1×10^{-8} g Ag/L of air in a 40-hour week. What is the allowed exposure in terms of atoms of Ag/L/week?

33. Determine

(a) the number of atoms in 0.185 g of palladium (Pd).
(b) the mass of 127 protons of palladium.

34. For bismuth (Bi), determine

(a) the number of atoms in 0.243 g.
(b) the mass of 139 protons.

35. The isotope Si-28 has a mass of 27.977 amu. For ten grams of Si-28, calculate

(a) the number of atoms.
(b) the total number of protons, neutrons, and electrons.

36. Myocardial perfusion imaging (MPI) is the latest tool used in the detection of coronary artery disease. It uses thallium-201 chloride. How many neutrons are in

(a) seventy atoms of thallium-201?
(b) one picogram (1×10^{-12} g) of thallium-201?

37. A cube of sodium has length 1.25 in. How many atoms are in that cube? (Note: $d_{Na} = 0.968$ g/cm^3.)

38. A cylindrical piece of pure copper ($d = 8.92$ g/cm^3) has diameter 1.15 cm and height 4.00 inches. How many atoms are in that cylinder? (Note: the volume of a right circular cylinder of radius r and height h is $V = \pi r^2 h$.)

2-4 Introduction to the Periodic Table

Elements and the Periodic Table

39. Give the symbols for

(a) potassium (b) cadmium (c) aluminum
(d) antimony (e) phosphorus

40. Name the elements represented by

(a) S (b) Sc (c) Se (d) Si (e) Sr

41. Classify the elements in Question 39 as metals (main group, transition, or post-transition), nonmetals, or metalloids.

42. Classify the elements in Question 40 as metals (main group, transition, or post-transition), nonmetals, or metalloids.

43. How many metals are in the following groups?

(a) Group 1 (b) Group 13 (c) Group 17

44. How many nonmetals are in the following periods?

(a) period 2 (b) period 4 (c) period 6

45. Which group in the periodic table

(a) has one metalloid and no nonmetals?
(b) has no nonmetals or transition metals?
(c) has no metals or metalloids?

46. Which period of the periodic table

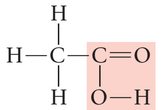
(a) has no metals?
(b) has no nonmetals?
(c) has one post-transition metal and two metalloids?

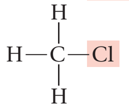
2-5 Molecules and Ions

47. Given the following condensed formulas, write the molecular formulas for the following molecules.

(a) dimethylamine $(CH_3)_2NH$
(b) propyl alcohol $CH_3(CH_2)_2OH$

48. Write the condensed structural formulas and molecular formulas for the following molecules. The reactive groups are shown in red.

(a)  (acetic acid)

(b)  (methyl chloride)

49. Give the number of protons and electrons in

(a) an N_2 molecule (identified in 1772).
(b) an N_3^- unit (synthesized in 1890).
(c) an N_5^+ unit (synthesized in 1999).
(d) an N_5N_5 salt (a U.S. Air Force research team's synthesis project).

50. Give the number of protons and electrons in the following:

(a) S_8 molecule. (b) SO_4^{2-} ion.
(c) H_2S molecule. (d) S^{2-} ion.

51. Complete the table given below. If necessary, use the periodic table.

Nuclear Symbol	Charge	Number of Protons	Number of Neutrons	Number of Electrons
_____	0	9	10	_____
^{31}P	0	_____	16	_____
_____	+3	27	30	_____
_____	_____	16	16	18

52. Complete the table given below.

Nuclear Symbol	Metal, Nonmetal, Metalloid	Group	Period	Number of Neutrons
_____	_____	13	3	14
_____	metalloid	16	_____	76
Xe-134	_____	_____	_____	_____
_____	nonmetal	14	_____	8

53. Classify the following compounds as electrolytes or nonelectrolytes.
(a) potassium chloride, KCl
(b) hydrogen peroxide, H_2O_2
(c) methane, CH_4
(d) barium nitrate, $Ba(NO_3)_2$

54. Which (if any) of the following compounds are nonelectrolytes?
(a) citric acid ($C_6H_8O_7$)
(b) calcium nitrate, $Ca(NO_3)_2$
(c) ammonium carbonate, $(NH_4)_2CO_3$
(d) iodine tribromide (IBr_3)

2-6 Formulas of Ionic Compounds
2-7 Names of Compounds

Names and Formulas of Ionic and Molecular Compounds

55. Write the formulas for the following molecules.
(a) methane (b) carbon tetraiodide
(c) hydrogen peroxide (d) nitrogen oxide
(e) silicon dioxide

56. Write the formulas of the following molecules
(a) silicon disulfide
(b) phosphine
(c) diphosphorus pentoxide
(d) radon tetrafluoride
(e) nitrogen trichloride

57. Write the names of the following molecules.
(a) ICl_3 (b) N_2O_5 (c) PH_3 (d) CBr_4 (e) SO_3

58. Write the names of the following molecules.
(a) $TeCl_2$ (b) I_4O_9 (c) XeO_3 (d) S_4N_4

59. Give the formulas of all the compounds containing no ions other than K^+, Ca^{2+}, Cl^-, and S^{2-}.

60. Give the formulas of compounds in which
(a) the cation is Ba^{2+}, the anion is I^- or N^{3-}.
(b) the anion is O^{2-}, the cation is Fe^{2+} or Fe^{3+}.

61. Write the formulas of the following ionic compounds.
(a) iron(III) acetate
(b) calcium nitrate
(c) potassium oxide
(d) gold(III) chloride
(e) barium nitride

62. Write formulas for the following ionic compounds:
(a) potassium hydrogen phosphate
(b) magnesium nitride
(c) lead(IV) bromide
(d) scandium(III) chloride
(e) barium acetate

63. Write the names of the following ionic compounds.
(a) $K_2Cr_2O_7$ (b) $Cu_3(PO_4)_2$ (c) $Ba(C_2H_3O_2)_2$
(d) AlN (e) $Co(NO_3)_2$

64. Write the names of the following ionic compounds.
(a) $Pb(C_2H_3O_2)_2$
(b) $Al(OH)_3$
(c) $Sr(HCO_3)_2$
(d) $Cu_3(PO_4)_2$
(e) Rb_2O

65. Write the names of the following ionic compounds.
(a) $HCl(aq)$ (b) $HClO_3(aq)$ (c) $Fe_2(SO_3)_3$
(d) $Ba(NO_2)_2$ (e) NaClO

66. Write formulas for the following ionic compounds.
(a) nitric acid (b) potassium sulfate
(c) iron(III) perchlorate (d) aluminum iodate
(e) sulfurous acid

67. Complete the following table.

Name	Formula
nitrous acid	_____
_____	$Ni(IO_3)_2$
gold(III) sulfide	_____
_____	$H_2SO_3(aq)$
nitrogen trifluoride	_____

68. Complete the following table.

Name	Formula
bromous acid	_____
_____	$Co(IO_4)_3$
platinum(IV) chloride	_____
_____	$HNO_2(aq)$
dinitrogen trioxide	_____

Unclassified

69. Write the formulas and names of the following:
(a) An ionic compound whose cation is a transition metal with 25 protons and 22 electrons and whose anion is an oxoanion of nitrogen with two oxygen atoms.
(b) A molecule made up of a metalloid in Group 13 and three atoms of a halogen in period 2.
(c) An ionic compound made up of an alkaline earth with 20 protons, and an anion with one hydrogen atom, a carbon atom, and 3 oxygen atoms.

70. Identify the following elements:
(a) A member of the same period as selenium but with two fewer protons than selenium.
(b) A transition metal in group 6, period 6.
(c) An alkaline earth with 38 protons.
(d) A post-transition metal in group 15.

71. Identify the following elements.
(a) A member of the same period as tellurium but with three fewer protons than tellurium.
(b) A post-transition metal in Group 14
(c) An alkali with 19 protons
(d) A metalloid in Group 15 with more than 40 protons

72. A molecule of ethylamine is made up of two carbon atoms, seven hydrogen atoms, and one nitrogen atom.
  (**a**) Write its molecular formula.
  (**b**) The reactive group in ethylamine is NH_2. Write its condensed structural formula.

73. Criticize each of the following statements:
  (**a**) The Rutherford experiment confirmed the presence of electrons in an atom.
  (**b**) A period is that segment of the periodic table in which elements are arranged by similar chemical properties.
  (**c**) Isotopes have the same mass number.
  (**d**) The molecule Be_3N_2 is triberylium dinitride.

74. Which of the following statements is/are always true? Never true? Usually true?
  (**a**) Compounds containing chlorine can be either molecular or ionic.
  (**b**) An ionic compound always has at least one metal.
  (**c**) When an element in a molecule has a "di" prefix, it means that the element has a +2 charge.

75. Some brands of salami contain 0.090% sodium benzoate ($NaC_7H_5O_2$) as a preservative. If you eat 6.00 oz of this salami, how many atoms of sodium will you consume, assuming salami contains no other source of that element?

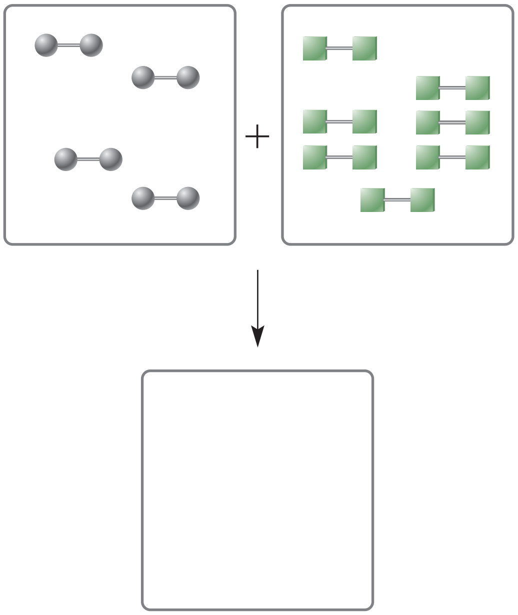
76. Carbon tetrachloride, CCl_4, was a popular dry-cleaning agent until it was shown to be carcinogenic. It has a density of 1.589 g/cm^3. What volume of carbon tetrachloride will contain a total of 6.00×10^{25} molecules of CCl_4?

Conceptual Problems

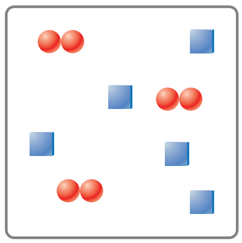
77. Which of the following statements are true?
  (**a**) Neutrons have neither mass nor charge.
  (**b**) Isotopes of an element have an identical number of protons.
  (**c**) C-14 and N-14 have identical neutron/proton (n/p^+) ratios.
  (**d**) The vertical columns in a periodic table are referred to as "groups."
  (**e**) When an atom loses an electron, it becomes positively charged.

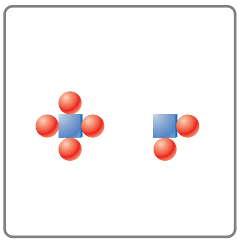
78. A student saw the following nuclear symbol for an unknown element: $^{23}_{11}X$. Which of the following statements about X and $^{23}_{11}X$ are true?
  (**a**) X is sodium.
  (**b**) X is vanadium.
  (**c**) X has 23 neutrons in its nucleus.
  (**d**) X^{2+} has 13 electrons.
  (**e**) $^{23}_{11}X$ has a proton/neutron ratio of about 1.1.

79. Using the laws of constant composition and the conservation of mass, complete the molecular picture of hydrogen molecules (○—○) reacting with chlorine molecules (□—□) to give hydrogen chloride (□—○) molecules.

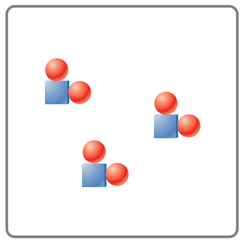


80. Use the law of conservation of mass to determine which numbered box(es) represent(s) the product mixture after the substances in the unnumbered box undergo a reaction.

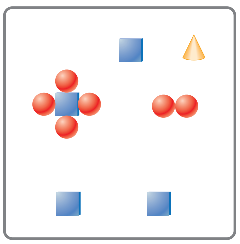




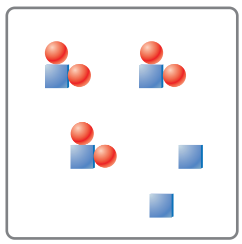
(1)



(2)



(3)



(4)

81. If squares represent carbon and spheres represent chlorine, make a representation of liquid CCl_4.

82. If a square represents a sulfur atom and a circle represents a sodium atom, sketch a representation of a sodium sulfide crystal.

83. Scientists are trying to synthesize elements with more than 114 protons. State the expected atomic number of

(**a**) the newest inert gas.
(**b**) the new element with properties similar to those of the alkaline earth metals.
(**c**) the new element that will behave like the halogens.
(**d**) the new (nontransition) metal whose ion will have a +2 charge.
(**e**) the new element that will start period 8.

84. Write the nuclear symbol for the element whose mass number is 234 and has 60% more neutrons than protons.

85. Mercury(II) oxide, a red powder, can be decomposed by heating to produce liquid mercury and oxygen gas. When a sample of this compound is decomposed, 3.87 g of oxygen and 48.43 g of mercury are produced. In a second experiment, 15.68 g of mercury is allowed to react with an excess of oxygen and 16.93 g of red mercury(II) oxide is produced. Show that these results are consistent with the law of constant composition.

86. Write the atomic symbol for the element whose ion has a −2 charge, has 20 more neutrons than electrons, and has a mass number of 126.

87. Consider the elements oxygen, fluorine, argon, sulfur, potassium, and strontium. From this group of elements, which ones fit the descriptions below?

(**a**) Two elements that are metals
(**b**) Four elements that are nonmetals
(**c**) Three elements that are solid at room temperature
(**d**) An element that is found in nature as X_8
(**e**) One pair of elements that may form a molecular compound
(**f**) One pair of elements that may form an ionic compound with formula AX
(**g**) One pair of elements that may form an ionic compound with formula AX_2
(**h**) One pair of elements that may form an ionic compound with formula A_2X
(**i**) An element that can form no compounds
(**j**) Three elements that are gases at room temperature

Challenge Problems

88. Three compounds containing only carbon and hydrogen are analyzed. The results for the analysis of the first two compounds are given below:

Compound	Mass of Carbon (g)	Mass of Hydrogen (g)
A	28.5	2.39
B	34.7	11.6
C	16.2	—

Which, if any, of the following results for the mass of hydrogen in compound C follows the law of multiple proportions?

(**a**) 5.84 g (**b**) 3.47 g (**c**) 2.72 g

89. Ethane and ethylene are two gases containing only hydrogen and carbon atoms. In a certain sample of ethane, 4.53 g of hydrogen is combined with 18.0 g of carbon. In a sample of ethylene, 7.25 g of hydrogen is combined with 43.20 g of carbon.

(**a**) Show how the data illustrate the law of multiple proportions.
(**b**) Suggest reasonable formulas for the two compounds.

90. Calculate the average density of a single Al-27 atom by assuming that it is a sphere with a radius of 0.143 nm. The masses of a proton, electron, and neutron are 1.6726×10^{-24} g, 9.1094×10^{-28} g, and 1.6749×10^{-24} g, respectively. The volume of a sphere is $4\pi r^3/3$, where r is its radius. Express the answer in grams per cubic centimeter. The density of aluminum is found experimentally to be 2.70 g/cm^3. What does that suggest about the packing of aluminum atoms in the metal?

91. The mass of an atom of N-14 is 2.3440×10^{-23} g. Using that fact and other information in this chapter, find the mass of an N-14^{3-} ion.

92. Each time you inhale, you take in about 500 mL (two significant figures) of air, each milliliter of which contains 2.5×10^{19} molecules. In delivering the Gettysburg Address, Abraham Lincoln is estimated to have inhaled about 200 times.

(**a**) How many molecules did Lincoln take in?
(**b**) In the entire atmosphere, there are about 1.1×10^{44} molecules. What fraction of the molecules in the earth's atmosphere was inhaled by Lincoln at Gettysburg?
(**c**) In the next breath that you take, how many molecules were inhaled by Lincoln at Gettysburg?

93. Hydrogen gas is prepared in a lab experiment. In this experiment, 18.00 g of aluminum metal are mixed with 25 mL of HCl (d = 1.025 g/cm^3). After the experiment, there are 12.00 g of aluminum and 30.95 g of a solution made up of water, aluminum ions, and chloride ions. Assuming no loss of products, how many liters of H_2 gas are obtained? The density of hydrogen gas at the temperature and pressure of the experiment is 0.0824 g/L.