2 Chemical Formulas and Composition Stoichiometry

OUTLINE

OBJECTIVES

After you have studied this chapter, you should be able to

- *Understand some early concepts of atoms*
- *Use chemical formulas to solve various kinds of chemical problems*
- *Relate names to formulas and charges of simple ions*
- *Combine simple ions to write formulas and names of some ionic compounds*
- *Recognize and use formula weights and mole relationships*
- *Interconvert masses, moles, and formulas*
- *Determine percent compositions in compounds*
- *Determine formulas from composition*
- *Perform calculations of purity of substances*

Some minerals and gems, **left to right:** ***Galena, PbS; aquamarine, $Be_3AlSi_6O_{18}$ (colored by trace amounts of iron); ruby, Al_2O_3 (with some Cr^{3+} ions replacing Al^{3+} ions); sulfur, S_8; silver, Ag; elbaite, $Na(Li,Al)_3Al_6Si_6O_{18}(BO_3)_3(OH)_4$ (the notation [Li,Al] indicates a variable ratio of lithium and aluminum among elbaite samples); sapphire, Al_2O_3 (with various colors due to different metal ions substituting for Al^{3+} ions); fluorite, CaF_2; copper, Cu; azurite, $Cu_3(CO_3)_2(OH)_2$; and malachite, $Cu_2CO_3(OH)_2$.***

The language we use to describe the forms of matter and the changes in its composition is not limited to use in chemistry courses; it appears throughout the scientific world. Chemical symbols, formulas, and equations are used in such diverse areas as agriculture, home economics, engineering, geology, physics, biology, medicine, and dentistry. In this chapter we describe the simplest atomic theory. We shall use it as we represent the chemical formulas of elements and compounds. Later, after additional facts have been introduced, this theory will be expanded.

It is important to learn this fundamental material well so that you can use it correctly and effectively.

The word "stoichiometry" is derived from the Greek *stoicheion*, which means "first principle or element," and *metron*, which means "measure." **Stoichiometry** describes the quantitative relationships among elements in compounds (composition stoichiometry) and among substances as they undergo chemical changes (reaction stoichiometry). In this chapter we are concerned with chemical formulas and composition stoichiometry. In Chapter 3 we shall discuss chemical equations and reaction stoichiometry.

2-1 ATOMS AND MOLECULES

The term "atom" comes from the Greek language and means "not divided" or "indivisible."

The Greek philosopher Democritus (470–400 BC) suggested that all matter is composed of tiny, discrete, indivisible particles that he called atoms. His ideas, based entirely on philosophical speculation rather than experimental evidence, were rejected for 2000 years. By the late 1700s, scientists began to realize that the concept of atoms provided an explanation for many experimental observations about the nature of matter.

By the early 1800s, the Law of Conservation of Matter (Section 1-1) and the Law of Definite Proportions (Section 1-5) were both accepted as general descriptions of how matter behaves. John Dalton (1766–1844), an English schoolteacher, tried to explain why matter behaves in such systematic ways as those expressed here. In 1808, he published the first "modern" ideas about the existence and nature of atoms. Dalton's explanation summarized and expanded the nebulous concepts of early philosophers and scientists; more importantly, his ideas were based on *reproducible experimental results* of measurements by many scientists. These ideas form the core of **Dalton's Atomic Theory,** one of the highlights in the history of scientific thought. In condensed form, Dalton's ideas may be stated as follows:

The radius of a calcium atom is only 1.97×10^{-8} cm, and its mass is 6.66×10^{-23}g.

Statement 3 is true for *chemical* reactions. It is not true, however, for *nuclear* reactions (Chapter 26).

1. An element is composed of extremely small, indivisible particles called atoms.
2. All atoms of a given element have identical properties that differ from those of other elements.
3. Atoms cannot be created, destroyed, or transformed into atoms of another element.
4. Compounds are formed when atoms of different elements combine with one another in small whole-number ratios.
5. The relative numbers and kinds of atoms are constant in a given compound.

Dalton believed that atoms were solid, indivisible spheres, an idea we now reject. But he showed remarkable insight into the nature of matter and its interactions. Some of his ideas could not be verified (or refuted) experimentally at the time. They were based on the limited experimental observations of his day. Even with their shortcomings, Dalton's ideas provided a framework that could be modified and expanded by later scientists. Thus John Dalton is often considered to be the father of modern atomic theory.

The smallest particle of an element that maintains its chemical identity through all chemical and physical changes is called an **atom** (Figure 2-1). In Chapter 5, we shall study the structure of the atom in detail; let us simply summarize here the main features of atomic composition. Atoms, and therefore *all* matter, consist principally of three **fundamental particles:** *electrons*, *protons*, and *neutrons*. These are the basic building blocks of

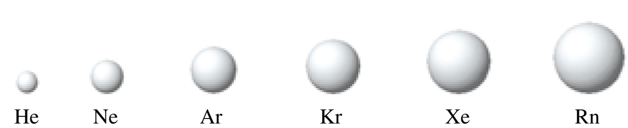


Figure 2-1 Relative sizes of monatomic molecules (single atoms) of the noble gases.

TABLE 2-1 *Fundamental Particles of Matter*

Particle (symbol)	Approximate Mass (amu)*	Charge (relative scale)
electron (e^-)	0.0	1−
proton (p or p^+)	1.0	1+
neutron (n or n^0)	1.0	none

*$1\ amu = 1.6605 \times 10^{-24}\ g$

atoms. The masses and charges of the three fundamental particles are shown in Table 2-1. The masses of protons and neutrons are nearly equal, but the mass of an electron is much smaller. Neutrons carry no charge. The charge on a proton is equal in magnitude, but opposite in sign, to the charge on an electron. Because atoms are electrically neutral,

an atom contains equal numbers of electrons and protons.

The **atomic number** (symbol is Z) of an element is defined as the number of protons in the nucleus. In the periodic table, elements are arranged in order of increasing atomic numbers. These are the red numbers above the symbols for the elements in the periodic table on the inside front cover. For example, the atomic number of silver is 47.

47 ← atomic number
Ag ← symbol

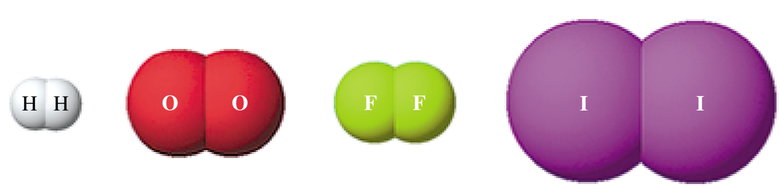
A **molecule** is the smallest particle of an element or compound that can have a stable independent existence. In nearly all molecules, two or more atoms are bonded together in very small, discrete units (particles) that are electrically neutral.

For Group VIIIA elements, the noble gases, a molecule contains only one atom, and so an atom and a molecule are the same (see Figure 2-1).

Individual oxygen atoms are not stable at room temperature and atmospheric pressure. Single atoms of oxygen mixed under these conditions quickly combine to form pairs. The oxygen with which we are all familiar is made up of two atoms of oxygen; it is a *diatomic* molecule, O_2. Hydrogen, nitrogen, fluorine, chlorine, bromine, and iodine are other examples of diatomic molecules (Figure 2-2).

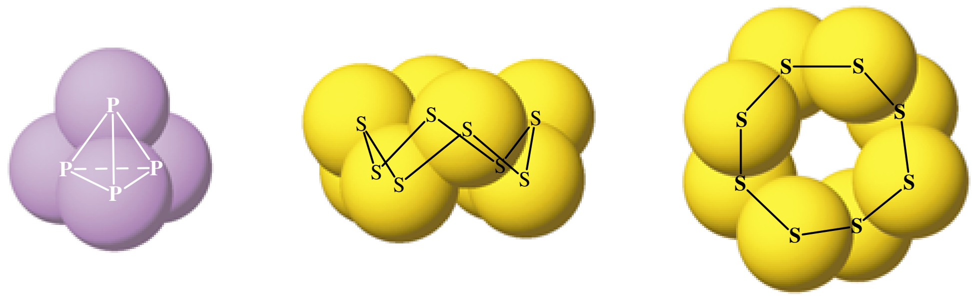
You should remember the common elements that occur as diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2.

Some other elements exist as more complex molecules. One form of phosphorus molecules consists of four atoms, and sulfur exists as eight-atom molecules at ordinary temperatures and pressures. Molecules that contain two or more atoms are called *polyatomic* molecules (Figure 2-3).



H_2 (hydrogen) O_2 (oxygen) F_2 (fluorine) I_2 (iodine)

Figure 2-2 Models of diatomic molecules of some elements, approximately to scale.



(a) (b) (c)

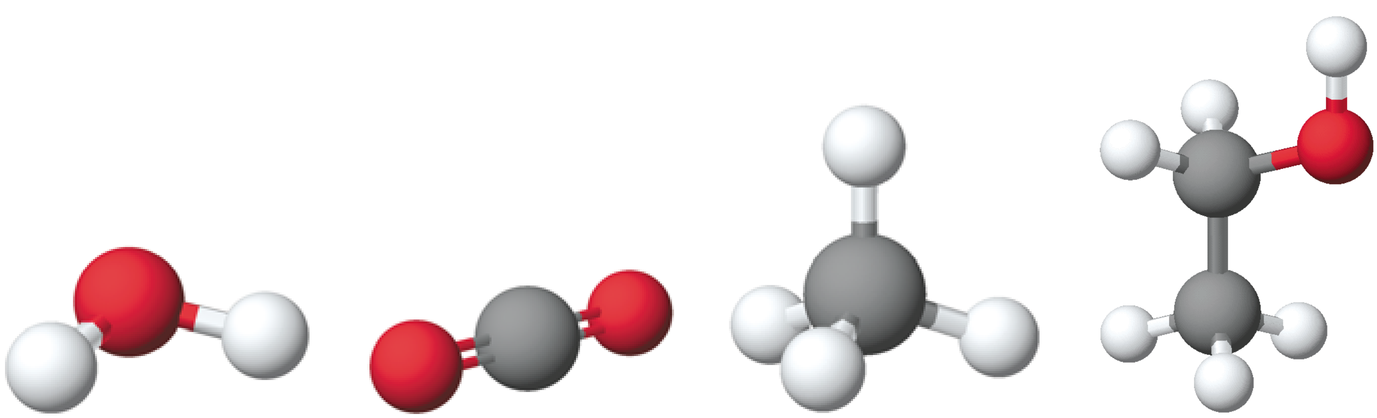
Figure 2-3 (a) A model of the P_4 molecule of white phosphorus. (b) A model of the S_8 ring found in rhombic sulfur. (c) Top view of the S_8 ring in rhombic sulfur.

In modern terminology, O_2 is named dioxygen, H_2 is dihydrogen, P_4 is tetraphosphorus, and so on. Even though such terminology is officially preferred, it has not yet gained wide acceptance. Most chemists still refer to O_2 as oxygen, H_2 as hydrogen, P_4 as phosphorus, and so on.

Methane is the principal component of natural gas.

Molecules of compounds are composed of more than one kind of atom. A water molecule consists of two atoms of hydrogen and one atom of oxygen. A molecule of methane consists of one carbon atom and four hydrogen atoms. The shapes of a few molecules are shown in Figure 2-4.

Atoms are the components of molecules, and molecules are the stable forms of many elements and compounds. We are able to study samples of compounds and elements that consist of large numbers of atoms and molecules. With the scanning tunnelling microscope it is now possible to "see" atoms (Figure 2-5). It would take millions of atoms to make a row as long as the diameter of the period at the end of this sentence.



H_2O (water) CO_2 (carbon dioxide) CH_4 (methane) C_2H_5OH (ethyl alcohol)

Figure 2-4 Formulas and models for molecules of some compounds.

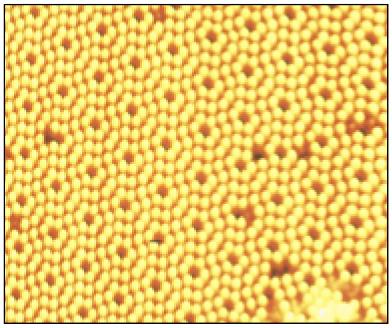
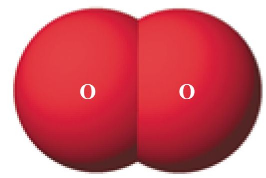


Figure 2-5 A computer reconstruction of the surface of a sample of silicon, as observed with a scanning tunnelling electron microscope (STM), reveals the regular pattern of individual silicon atoms. Many important reactions occur on the surfaces of solids. Observations of the atomic arrangements on surfaces help chemists understand such reactions. New information available using the STM will give many details about chemical bonding in solids.

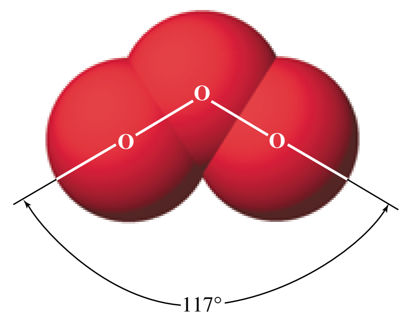
2-2 CHEMICAL FORMULAS

The **chemical formula** for a substance shows its chemical composition. This represents the elements present as well as the ratio in which the atoms of the elements occur. The formula for a single atom is the same as the symbol for the element. Thus, Na can represent a single sodium atom. It is unusual to find such isolated atoms in nature, with the exception of the noble gases (He, Ne, Ar, Kr, Xe, and Rn). A subscript following the symbol of an element indicates the number of atoms in a molecule. For instance, F_2 indicates a molecule containing two fluorine atoms, and P_4 a molecule containing four phosphorus atoms.



An O_2 molecule.

Some elements exist in more than one form. Familiar examples include (1) oxygen, found as O_2 molecules, and ozone, found as O_3 molecules, and (2) two crystalline forms of carbon—diamond and graphite (Figure 13-33). Different forms of the same element in the same physical state are called **allotropic modifications,** or **allotropes.**



An O_3 molecule.

Compounds contain two or more elements in chemical combination in fixed proportions. Many compounds exist as molecules (Table 2-2). Hence, each molecule of hydrogen chloride, HCl, contains one atom of hydrogen and one atom of chlorine; each molecule of carbon tetrachloride, CCl_4, contains one carbon atom and four chlorine atoms.

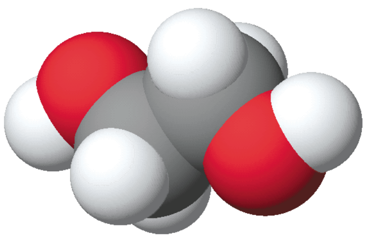
TABLE 2-2 *Names and Formulas of Some Common Molecular Compounds*

Name	Formula	Name	Formula	Name	Formula
water	H_2O	sulfur dioxide	SO_2	butane	C_4H_{10}
hydrogen peroxide	H_2O_2	sulfur trioxide	SO_3	pentane	C_5H_{12}
hydrogen chloride*	HCl	carbon monoxide	CO	benzene	C_6H_6
sulfuric acid	H_2SO_4	carbon dioxide	CO_2	methanol (methyl alcohol)	CH_3OH
nitric acid	HNO_3	methane	CH_4	ethanol (ethyl alcohol)	CH_3CH_2OH
acetic acid	CH_3COOH	ethane	C_2H_6	acetone	CH_3COCH_3
ammonia	NH_3	propane	C_3H_8	diethyl ether (ether)	$CH_3CH_2—O—CH_2CH_3$

**Called hydrochloric acid if dissolved in water.*

An aspirin molecule, $C_9H_8O_4$, contains nine carbon atoms, eight hydrogen atoms, and four oxygen atoms.

Many of the molecules found in nature are organic compounds. **Organic compounds** contain C—C or C—H bonds or both. Eleven of the compounds listed in Table 2-2 are organic compounds (acetic acid and the last ten entries). All of the other compounds in the table are **inorganic compounds.**



A model of ethylene glycol.

Some groups of atoms behave chemically as single entities. For instance, an oxygen atom that is bonded to a hydrogen atom and also to a carbon atom that is bonded to three other atoms forms the reactive combination of atoms known as the alcohol group or molecule. In formulas of compounds containing two or more of the same group, the group formula is enclosed in parentheses. Thus, ethylene glycol contains two *alcohol groups* and its formula is $C_2H_4(OH)_2$ (see structure in the margin). When you count the number of atoms in this molecule from its formula, you must multiply the numbers of hydrogen and oxygen atoms in the OH group by 2. There are *two* carbon atoms, *six* hydrogen atoms and *two* oxygen atoms in a molecule of ethylene glycol.

Compounds were first recognized as distinct substances because of their different physical properties and because they could be separated from one another by physical methods. Once the concept of atoms and molecules was established, the reason for these differences in properties could be understood: Two compounds differ from each other because their molecules are different. Conversely, if two molecules contain the same number of the same kinds of atoms, arranged the same way, then both are molecules of the same compound. Thus, the atomic theory explains the **Law of Definite Proportions** (see Section 1-5).

This law, also known as the **Law of Constant Composition,** can now be extended to include its interpretation in terms of atoms. It is so important for performing the calculation in this chapter that we restate it here:

> Different pure samples of a compound always contain the same elements in the same proportion by mass; this corresponds to atoms of these elements combined in fixed numerical ratios.

So we see that for a substance composed of molecules, the **chemical formula** gives the number of atoms of each type in the molecule. But this formula does not express the order in which the atoms in the molecules are bonded together. The **structural formula** shows the order in which atoms are connected. The lines connecting atomic symbols represent chemical bonds between atoms. The bonds are actually forces that tend to hold atoms at certain distances and angles from one another. For instance, the structural formula of propane shows that the three C atoms are linked in a chain, with three H atoms bonded to each of the end C atoms and two H atoms bonded to the center C. **Ball-and-stick** molecular models and **space-filling** molecular models help us to see the shapes and relative sizes of molecules. These four representations are shown in Figure 2-6. The ball-and-stick and space-filling models show (1) the *bonding sequence*, that is the order in which the atoms are connected to each other, and (2) the *geometrical arrangements* of the atoms in the molecule. As we shall see later, both are extremely important because they determine the properties of compounds.

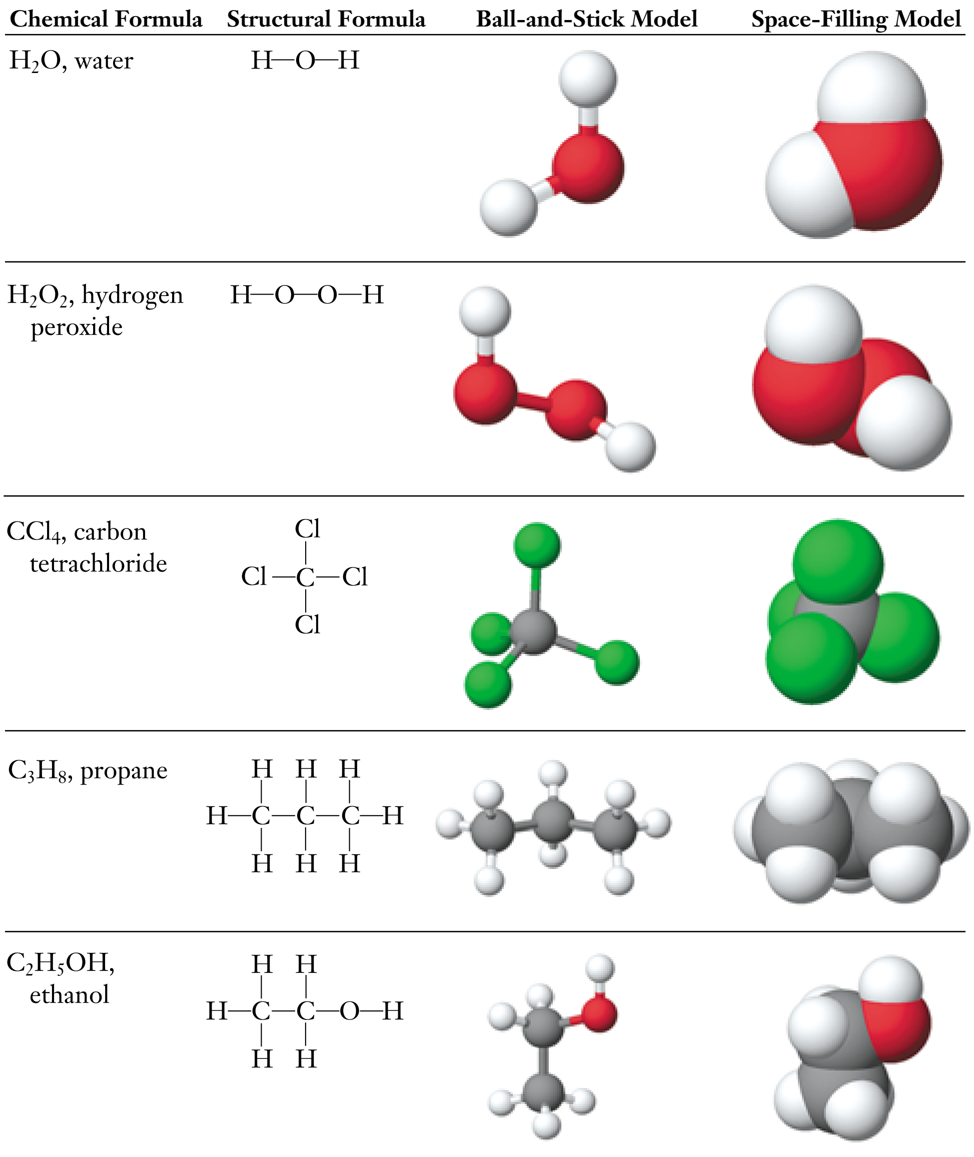


Figure 2-6 Formulas and models for some molecules. Structural formulas show the order in which atoms are connected but do not represent true molecular shapes. Ball-and-stick models use balls of different colors to represent atoms and sticks to represent bonds; they show the three-dimensional shapes of molecules. Space-filling models show the (approximate) relative sizes of atoms and the shapes of molecules.

2-3 IONS AND IONIC COMPOUNDS

So far we have discussed only compounds that exist as discrete molecules. Some compounds, such as sodium chloride, NaCl, consist of collections of large numbers of ions. An **ion** is an atom or group of atoms that carries an electric charge. Ions that possess a *positive* charge, such as the sodium ion, Na^+, are called **cations.** Those carrying a *negative* charge, such as the chloride ion, Cl^-, are called **anions.** The charge on an ion *must*

The words "cation" (kat′-i-on) and "anion" (an′-i-on) and their relationship to cathode and anode will be described in Chapter 21.

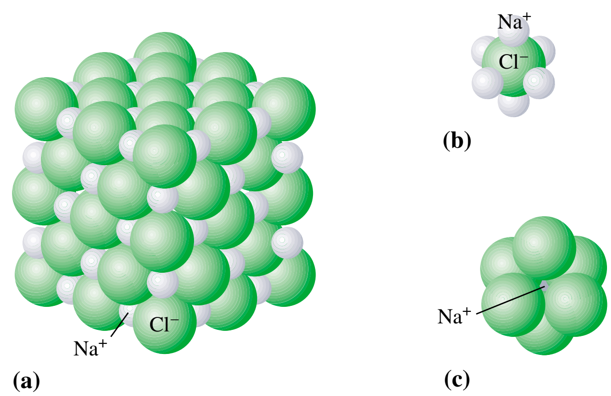


Figure 2-7 The arrangement of ions in NaCl. (a) A crystal of sodium chloride consists of an extended array that contains equal numbers of sodium ions *(small spheres)* and chloride ions *(large spheres)*. Within the crystal, (b) each chloride ion is surrounded by six sodium ions, and (c) each sodium ion is surrounded by six chloride ions.

be included as a superscript on the right side of the chemical symbol(s) when we write the formula for the individual ion.

As discussed in detail in Chapter 5, an atom consists of a very small, very dense, positively charged *nucleus* surrounded by a diffuse distribution of negatively charged particles called *electrons.* The number of positive charges in the nucleus defines the identity of the element to which the atom corresponds. Electrically neutral atoms contain the same number of electrons outside the nucleus as positive charges (protons) within the nucleus. Ions are formed when neutral atoms lose or gain electrons. An Na^+ ion is formed when a sodium atom loses one electron, and a Cl^- ion is formed when a chlorine atom gains one electron.

The compound NaCl consists of an extended array of Na^+ and Cl^- ions (Figure 2-7). Within the crystal (though not on the surface) each Na^+ ion is surrounded at equal distances by six Cl^- ions, and each Cl^- ion is similarly surrounded by six Na^+ ions. *Any* compound, whether ionic or molecular, is electrically neutral; that is, it has no net charge. In NaCl this means that the Na^+ and Cl^- ions are present in a 1 : 1 ratio, and this is indicated by the formula NaCl.

The general term "formula unit" applies to molecular or ionic compounds, whereas the more specific term "molecule" applies only to elements and compounds that exist as discrete molecules.

Because there are no "molecules" of ionic substances, we should not refer to "a molecule of sodium chloride, NaCl," for example. Instead, we refer to a **formula unit** of NaCl, which consists of one Na^+ ion and one Cl^- ion. Likewise, one formula unit of $CaCl_2$ consists of one Ca^{2+} ion and two Cl^- ions. As you will see in the next section, we speak of the formula unit of all ionic compounds as the smallest, whole-number ratios of ions that yield neutral representations. It is also acceptable to refer to a formula unit of a molecular compound. One formula unit of propane, C_3H_8, is the same as one molecule of C_3H_8; it contains three C atoms and eight H atoms bonded together into a group.

For the present, we shall tell you which substances are ionic and which are molecular when it is important to know. Later you will learn to make the distinction yourself.

In this text, we use the standard convention of representing multiple charges with the number before the sign, e.g., Ca^{2+}, *not* Ca^{+2} and $SO_4{}^{2-}$, *not* $SO_4{}^{-2}$.

Polyatomic ions are groups of atoms that bear an electric charge. Examples include the ammonium ion, $NH_4{}^+$; the sulfate ion, $SO_4{}^{2-}$; and the nitrate ion, $NO_3{}^-$. Table 2-3 shows the formulas, ionic charges, and names of some common ions. When writing the formula of a polyatomic compound, we show groups in parentheses when they appear more than once. For example, $(NH_4)_2SO_4$ represents a compound that has two $NH_4{}^+$ ions for each $SO_4{}^{2-}$ ion.

TABLE 2-3 *Formulas, Ionic Charges, and Names of Some Common Ions*

Common Cations (positive ions)			Common Anions (negative ions)		
Formula	*Charge*	*Name*	*Formula*	*Charge*	*Name*
Na^+	1+	sodium	F^-	1−	fluoride
K^+	1+	potassium	Cl^-	1−	chloride
NH_4^+	1+	ammonium	Br^-	1−	bromide
Ag^+	1+	silver	OH^-	1−	hydroxide
			CH_3COO^-	1−	acetate
Mg^{2+}	2+	magnesium	NO_3^-	1−	nitrate
Ca^{2+}	2+	calcium			
Zn^{2+}	2+	zinc	O^{2-}	2−	oxide
Cu^+	1+	copper(I)	S^{2-}	2−	sulfide
Cu^{2+}	2+	copper(II)	SO_4^{2-}	2−	sulfate
Fe^{2+}	2+	iron(II)	SO_3^{2-}	2−	sulfite
			CO_3^{2-}	2−	carbonate
Fe^{3+}	3+	iron(III)			
Al^{3+}	3+	aluminum	PO_4^{3-}	3−	phosphate

As we shall see, some metals can form more than one kind of ion with a positive charge. For such metals, we specify which ion we mean with a Roman numeral—e.g., iron(II) or iron(III). Because zinc forms no stable ions other than Zn^{2+}, we do not need to use Roman numerals in its name.

2-4 NAMES AND FORMULAS OF SOME IONIC COMPOUNDS

During your study of chemistry you will have many occasions to refer to compounds by name. In this section, we shall see how a few compounds should be named. More comprehensive rules for naming compounds are presented at the appropriate places later in the text.

Table 2-2 includes examples of names for a few common molecular compounds. You should learn that short list before proceeding much farther in this textbook. We shall name many more molecular compounds as we encounter them in later chapters.

Ionic compounds *(clockwise, from top):* salt (sodium chloride, NaCl), calcite (calcium carbonate, $CaCO_3$), cobalt(II) chloride hexahydrate, ($CoCl_2 \cdot 6H_2O$), fluorite (calcium fluoride, CaF_2).

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 3.13, Naming Ionic Compounds.

The names of some common ions appear in Table 2-3. You will need to know the names and formulas of these frequently encountered ions. They can be used to write the formulas and names of many ionic compounds. We write the formula of an ionic compound by adjusting the relative numbers of positive and negative ions so their total charges cancel (i.e., add to zero). The name of an ionic compound is formed by giving the names of the ions, with the positive ion named first.

Problem-Solving Tip: ***Where to Start in Learning to Name Compounds***

You may not be sure of the best point to start learning the naming of compounds. It has been found that before rules for naming can make much sense or before we can expand our knowledge to more complex compounds, we need to know the names and formulas in Tables 2-2 and 2-3. If you are unsure of your ability to recall a name or a formula in Tables 2-2 and 2-3 when given the other, prepare flash cards, lists, and so on that you can use to learn these tables.

EXAMPLE 2-1 *Formulas for Ionic Compounds*

Write the formulas for the following ionic compounds: (a) sodium fluoride, (b) calcium fluoride, (c) iron(II) sulfate, (d) zinc phosphate.

Plan

In each case, we identify the chemical formulas of the ions from Table 2-3. These ions must be present in the simplest whole-number ratio that gives the compound *no net charge.* Recall that the formulas and names of ionic compounds are written by giving the positively charged ion first.

Solution

(a) The formula for the sodium ion is Na^+, and the formula for the fluoride ion is F^- (Table 2-3). Because the charges on these two ions are equal in magnitude, the ions must be present in equal numbers, or in a 1:1 ratio. Thus, the formula for sodium fluoride is NaF.

(b) The formula for the calcium ion is Ca^{2+} and the formula for the fluoride ion is F^-. Now each positive ion (Ca^{2+}) provides twice as much charge as each negative ion (F^-). So there must be twice as many F^- ions as Ca^{2+} ions to equalize the charge. This means that the ratio of calcium to fluoride ions is 1:2. So the formula for calcium fluoride is CaF_2.

(c) The iron(II) ion is Fe^{2+}, and the sulfate ion is $SO_4{}^{2-}$. As in part (a), the equal magnitudes of positive and negative charges tell us that the ions must be present in equal numbers, or in a 1:1 ratio. The formula for iron(II) sulfate is $FeSO_4$.

(d) The zinc ion is Zn^{2+}, and the phosphate ion is $PO_4{}^{3-}$. Now it will take *three* Zn^{2+} ions to account for as much charge (6+ total) as would be present in *two* $PO_4{}^{3-}$ ions (6− total). So the formula for zinc phosphate is $Zn_3(PO_4)_2$.

You should now work Exercises 14 and 21.

EXAMPLE 2-2 *Names for Ionic Compounds*

Name the following ionic compounds: (a) $(NH_4)_2S$, (b) $Cu(NO_3)_2$, (c) $ZnCl_2$, (d) $Fe_2(CO_3)_3$.

Plan

In naming ionic compounds, it is helpful to inspect the formula for atoms or groups of atoms that we recognize as representing familiar ions.

Solution

(a) The presence of the polyatomic grouping NH_4 in the formula suggests to us the presence of the ammonium ion, NH_4^+. There are two of these, each accounting for 1+ in charge. To balance this, the single S must account for 2− in charge, or S^{2-}, which we recognize as the sulfide ion. Thus, the name of the compound is ammonium sulfide.

(b) The NO_3 grouping in the formula tells us that the nitrate ion, NO_3^-, is present. Two of these nitrate ions account for $2 \times 1- = 2-$ in negative charge. To balance this, copper must account for 2+ charge and be the copper(II) ion. The name of the compound is copper(II) nitrate.

(c) The positive ion present is zinc ion, Zn^{2+}, and the negative ion is chloride, Cl^-. The name of the compound is zinc chloride.

(d) Each CO_3 grouping in the formula must represent the carbonate ion, CO_3^{2-}. The presence of *three* such ions accounts for a total of 6− in negative charge, so there must be a total of 6+ present in positive charge to balance this. It takes *two* iron ions to provide this 6+, so each ion must have a charge of 3+ and be Fe^{3+}, the iron(III) ion, or ferric ion. The name of the compound is iron(III) carbonate.

We use the information that the carbonate ion has a 2− charge to find the charge on the iron ions. The total charges must add up to zero.

You should now work Exercises 13 and 20.

A more extensive discussion on naming compounds appears in Chapter 4.

2-5 ATOMIC WEIGHTS

As the chemists of the eighteenth and nineteenth centuries painstakingly sought information about the compositions of compounds and tried to systematize their knowledge, it became apparent that each element has a characteristic mass relative to every other element. Although these early scientists did not have the experimental means to measure the mass of each kind of atom, they succeeded in defining a *relative* scale of atomic masses.

An early observation was that carbon and hydrogen have relative atomic masses, also traditionally called **atomic weights (AW),** of approximately 12 and 1, respectively. Thousands of experiments on the compositions of compounds have resulted in the establishment of a scale of relative atomic weights based on the **atomic mass unit (amu),** which is defined as *exactly* $\frac{1}{12}$ *of the mass of an atom of a particular kind of carbon atom, called carbon-12.*

The term "atomic weight" is widely accepted because of its traditional use, although it is properly a mass rather than a weight. "Atomic mass" is often used.

On this scale, the atomic weight of hydrogen (H) is 1.00794 amu, that of sodium (Na) is 22.989768 amu, and that of magnesium (Mg) is 24.3050 amu. This tells us that Na atoms have nearly 23 times the mass of H atoms, and Mg atoms are about 24 times heavier than H atoms.

When you need values of atomic weights, consult the periodic table or the alphabetical listing of elements, both found on facing pages inside the front cover.

2-6 THE MOLE

Even the smallest bit of matter that can be handled reliably contains an enormous number of atoms. So we must deal with large numbers of atoms in any real situation, and some unit for conveniently describing a large number of atoms is desirable. The idea of using a unit to describe a particular number (amount) of objects has been around for a long time. You are already familiar with the dozen (12 items) and the gross (144 items).

"Mole" is derived from the Latin word *moles*, which means "a mass." "Molecule" is the diminutive form of this word and means "a small mass."

The SI unit for amount is the **mole,** abbreviated mol. It is *defined* as the amount of substance that contains as many entities (atoms, molecules, or other particles) as there are atoms in exactly 0.012 kg of pure carbon-12 atoms. Many experiments have refined the number, and the currently accepted value is

$$1 \text{ mole} = 6.0221367 \times 10^{23} \text{ particles}$$

This number, often rounded off to 6.022×10^{23}, is called **Avogadro's number** in honor of Amedeo Avogadro (1776–1856), whose contributions to chemistry are discussed in Section 12-8.

According to its definition, the mole unit refers to a fixed number of items, the identities of which must be specified. Just as we speak of a dozen eggs or a pair of aces, we refer to a mole of atoms or a mole of molecules (or a mole of ions, electrons, or other particles). We could even think about a mole of eggs, although the size of the required carton staggers the imagination! Helium exists as discrete He atoms, so one mole of helium consists of 6.022×10^{23} He *atoms.* Hydrogen commonly exists as diatomic (two-atom) molecules, so one mole of hydrogen is 6.022×10^{23} H_2 *molecules* and $2(6.022 \times 10^{23})$ H atoms.

Every kind of atom, molecule, or ion has a definite characteristic mass. It follows that one mole of a given pure substance also has a definite mass, regardless of the source of the sample. This idea is of central importance in many calculations throughout the study of chemistry and the related sciences.

Because the mole is defined as the number of atoms in 0.012 kg (or 12 g) of carbon-12, and the atomic mass unit is defined as $\frac{1}{12}$ of the mass of a carbon-12 atom, the following convenient relationship is true:

> The mass of one mole of atoms of a pure element in grams is numerically equal to the atomic weight of that element in atomic mass units. This is also called the **molar mass** of the element; its units are grams/mole, also written as g/mol or $g{\cdot}mol^{-1}$.

For instance, if you obtain a pure sample of the metallic element titanium (Ti), whose atomic weight is 47.88 amu, and measure out 47.88 g of it, you will have one mole, or 6.022×10^{23} titanium atoms.

The symbol for an element can be used to (1) identify the element, (2) represent one atom of the element, or (3) represent one mole of atoms of the element. The last interpretation will be extremely useful in calculations in the next chapter.

The atomic weight of iron (Fe) is 55.847 amu. Suppose that one dozen large eggs weighs 24 oz.

A quantity of a substance may be expressed in a variety of ways. For example, consider a dozen eggs and 55.847 grams (or one mole) of iron (Figure 2-8). We can express the amount of eggs or iron present in any of several units. We can then construct unit factors to relate an amount of the substance expressed in one kind of unit to the same amount expressed in another unit.

Unit Factors for Eggs	Unit Factors for Iron
$\frac{12 \text{ eggs}}{1 \text{ doz eggs}}$	$\frac{6.022 \times 10^{23} \text{ Fe atoms}}{1 \text{ mol Fe atoms}}$
$\frac{12 \text{ eggs}}{24 \text{ oz eggs}}$	$\frac{6.022 \times 10^{23} \text{ Fe atoms}}{55.847 \text{ g Fe}}$
and so on	and so on

Figure 2-8 Three ways of representing amounts.

As Table 2-4 suggests, the concept of a mole as applied to atoms is especially useful. It provides a convenient basis for comparing the masses of equal numbers of atoms of different elements.

Figure 2-9 shows what one mole of atoms of each of some common elements looks like. Each of the examples in Figure 2-9 represents 6.022×10^{23} *atoms* of the element.

The relationship between the mass of a sample of an element and the number of moles of atoms in the sample is illustrated in Example 2-3.

In this textbook we usually work problems involving atomic weights (masses) or formula weights (masses) rounded to only one decimal place. We round the answer further if initial data do not support the number of significant figures obtained using the rounded atomic weights. Similarly, if the initial data indicate that more significant figures are justified, we will rework such problems using atomic weights and formula weights containing values beyond the tenths place.

TABLE 2-4 ***Mass of One Mole of Atoms of Some Common Elements***

Element	A Sample with a Mass of	Contains
carbon	12.0 g C	6.02×10^{23} C atoms or 1 mol of C atoms
titanium	47.9 g Ti	6.02×10^{23} Ti atoms or 1 mol of Ti atoms
gold	197.0 g Au	6.02×10^{23} Au atoms or 1 mol of Au atoms
hydrogen	1.0 g H_2	6.02×10^{23} H atoms or 1 mol of H atoms (3.01×10^{23} H_2 molecules or $\frac{1}{2}$ mol of H_2 molecules)
sulfur	32.1 g S_8	6.02×10^{23} S atoms or 1 mol of S atoms (0.753×10^{23} S_8 molecules or $\frac{1}{8}$ mol of S_8 molecules)

Figure 2-9 One mole of atoms of some common elements. Back row *(left to right)*: bromine, aluminum, mercury, copper. Front row *(left to right)*: sulfur, zinc, iron.

EXAMPLE 2-3 *Moles of Atoms*

How many moles of atoms does 136.9 g of iron metal contain?

Plan

To the required four significant figures, 1 mol Fe atoms = 55.85 g Fe.

The atomic weight of iron is 55.85 amu. This tells us that the molar mass of iron is 55.85 g/mol, or that one mole of iron atoms is 55.85 g of iron. We can express this as either of two unit factors:

$$\frac{1 \text{ mol Fe atoms}}{55.85 \text{ g Fe}} \quad \text{or} \quad \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe atoms}}$$

Because one mole of iron has a mass of 55.85 g, we expect that 136.9 g will be a fairly small number of moles (greater than 1, but less than 10).

Solution

$$\underline{?} \text{ mol Fe atoms} = 136.9 \text{ g Fe} \times \frac{1 \text{ mol Fe atoms}}{55.85 \text{ g Fe}} = \boxed{2.451 \text{ mol Fe atoms}}$$

You should now work Exercises 34 and 40.

Once the number of moles of atoms of an element is known, the number of atoms in the sample can be calculated, as Example 2-4 illustrates.

EXAMPLE 2-4 *Numbers of Atoms*

How many atoms are contained in 2.451 mol of iron?

Plan

One mole of atoms of an element contains Avogadro's number of atoms, or 6.022×10^{23} atoms. This lets us generate the two unit factors

$$\frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}} \quad \text{and} \quad \frac{1 \text{ mol atoms}}{6.022 \times 10^{23} \text{ atoms}}$$

Solution

$$\underline{?} \text{ Fe atoms} = 2.451 \text{ mol Fe atoms} \times \frac{6.022 \times 10^{23} \text{ Fe atoms}}{1 \text{ mol Fe atoms}} = \boxed{1.476 \times 10^{24} \text{ Fe atoms}}$$

We expected the number of atoms in more than two moles of atoms to be a very large number. Written in nonscientific notation, the answer to this example is: 1,476,000,000,000,000,000,000,000.

Try to name this number with its many zeroes.

You should now work Exercise 42.

If we know the atomic weight of an element on the carbon-12 scale, we can use the mole concept and Avogadro's number to calculate the *average* mass of one atom of that element in grams (or any other mass unit we choose).

CHEMISTRY IN USE

The Development of Science

Avogadro's Number

If you think that the value of Avogadro's number, 6×10^{23}, is too large to be useful to anyone but chemists, look up into the sky on a clear night. You may be able to see about 3000 stars with the naked eye, but the total number of stars swirling around you in the known universe is approximately equal to Avogadro's number. Just think, the known universe contains approximately one mole of stars! You don't have to leave earth to encounter such large numbers. The water in the Pacific Ocean has a volume of about 6×10^{23} mL and a mass of about 6×10^{23} g.

Avogadro's number is almost incomprehensibly large. For example, if one mole of dollars given away at the rate of a million per second beginning when the earth first formed some 4.5 billion years ago, would any remain today? Surprisingly, about three fourths of the original mole of dollars would be left today; it would take about 14,500,000,000 more years to give away the remaining money at $1 million per second.

Computers can be used to provide another illustration of the magnitude of Avogadro's number. If a computer can count up to one billion in one second, it would take that computer about 20 million years to count up to 6×10^{23}. In contrast, recorded human history goes back only a few thousand years.

The impressively large size of Avogadro's number can give us very important insights into the very small sizes of individual molecules. Suppose one drop of water evaporates in one hour. There are about 20 drops in one milliliter of water, which weighs one gram. So one drop of water is about 0.05 g of water. How many H_2O molecules evaporate per second?

$$\frac{\underline{?}\ H_2O \text{ molecules}}{1 \text{ s}} = \frac{0.05 \text{ g } H_2O}{1 \text{ h}} \times \frac{1 \text{ mol } H_2O}{18 \text{ g } H_2O} \times \frac{6 \times 10^{23}\ H_2O \text{ molecules}}{1 \text{ mol } H_2O} \times \frac{1 \text{ h}}{60 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ s}}$$
$$= 5 \times 10^{17}\ H_2O \text{ molecules/s}$$

5×10^{17} H_2O molecules evaporating per second is five hundred million billion H_2O molecules evaporating per second—a number that is beyond our comprehension! This calculation helps us to recognize that water molecules are incredibly small. There are approximately 1.7×10^{21} water molecules in a single drop of water.

By gaining some appreciation of the vastness of Avogadro's number, we gain a greater appreciation of the extremely tiny volumes occupied by individual atoms, molecules, and ions.

Ronald DeLorenzo
Middle Georgia College
Original concept by Larry Nordell

EXAMPLE 2-5 *Masses of Atoms*

Calculate the average mass of one iron atom in grams.

Plan

We expect that the mass of a single atom in grams would be a *very* small number. We know that one mole of Fe atoms has a mass of 55.85 g and contains 6.022×10^{23} Fe atoms. We use this information to generate unit factors to carry out the desired conversion.

Solution

To gain some appreciation of how little this is, write 9.274×10^{-23} gram as a decimal fraction, and try to name the fraction.

$$\frac{\underline{?}\text{ g Fe}}{\text{Fe atom}} = \frac{55.85\text{ g Fe}}{1\text{ mol Fe atoms}} \times \frac{1\text{ mol Fe atoms}}{6.022 \times 10^{23}\text{ Fe atoms}} = \boxed{9.274 \times 10^{-23}\text{ g Fe/Fe atom}}$$

Thus, we see that the average mass of one Fe atom is only 9.274×10^{-23} g, that is, 0.00000000000000000000009274 g.

Example 2-5 demonstrates how small atoms are and why it is necessary to use large numbers of atoms in practical work. The next example will help you to realize how large Avogadro's number is.

EXAMPLE 2-6 *Avogadro's Number*

A stack of 500 sheets of typing paper is 1.9 inches thick. Calculate the thickness, in inches and in miles, of a stack of typing paper that contains one mole (Avogadro's number) of sheets.

Plan

We construct unit factors from the data given, from conversion factors in Table 1-7, and from Avogadro's number.

Solution

$$\underline{?}\text{ in} = 1\text{ mol sheets} \times \frac{6.022 \times 10^{23}\text{ sheets}}{1\text{ mol sheets}} \times \frac{1.9\text{ in.}}{500\text{ sheets}} = \boxed{2.3 \times 10^{21}\text{ in.}}$$

$$\underline{?}\text{ mi} = 2.3 \times 10^{21}\text{ in.} \times \frac{1\text{ ft}}{12\text{ in.}} \times \frac{1\text{ mi}}{5280\text{ ft}} = \boxed{3.6 \times 10^{16}\text{ mi}}$$

Imagine the number of trees required to make this much paper!

By comparison, the sun is about 93 million miles from the earth. This stack of paper would make 390 million stacks that reach from the earth to the sun.

Problem-Solving Tip: *When Do We Round?*

Even though the number 1.9 has two significant figures, we carry the other numbers in Example 2-6 to more significant figures. Then we round at the end to the appropriate number of significant figures. The numbers in the distance conversions are exact numbers.

2-7 FORMULA WEIGHTS, MOLECULAR WEIGHTS, AND MOLES

The **formula weight (FW)** of a substance is the sum of the atomic weights (AW) of the elements in the formula, each taken the number of times the element occurs. Hence a formula weight gives the mass of one formula unit in atomic mass units.

Formula weights, like the atomic weights on which they are based, are relative masses. The formula weight for sodium hydroxide, NaOH, (rounded off to the nearest 0.1 amu) is found as follows.

Number of Atoms of Stated Kind		× Mass of One Atom	= Mass Due to Element
1 × Na =	1	× 23.0 amu	= 23.0 amu of Na
1 × H =	1	× 1.0 amu	= 1.0 amu of H
1 × O =	1	× 16.0 amu	= 16.0 amu of O
		Formula weight of NaOH	= 40.0 amu

The term "formula weight" is correctly used for either ionic or molecular substances. When we refer specifically to molecular (nonionic) substances, that is, substances that exist as discrete molecules, we often substitute the term **molecular weight (MW).**

EXAMPLE 2-7 *Formula Weights*

Calculate the formula weight (molecular weight) of acetic acid (vinegar), CH_3COOH, using rounded values for atomic weights given in the International Table of Atomic Weights inside the front cover of the text.

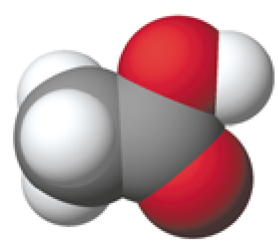
Plan

We add the atomic weights of the elements in the formula, each multiplied by the number of times the element occurs.

Solution

Number of Atoms of Stated Kind		× Mass of One Atom	= Mass Due to Element
2 × C =	2	× 12.0 amu	= 24.0 amu of C
4 × H =	4	× 1.0 amu	= 4.0 amu of H
2 × O =	2	× 16.0 amu	= 32.0 amu of O
		Formula weight (molecular weight) of acetic acid (vinegar)	= 60.0 amu

You should now work Exercise 26.



A space-filling model of an acetic acid (vinegar) molecule, CH_3COOH.

TABLE 2-5 *One Mole of Some Common Molecular Substances*

Substance	Molecular Weight	A Sample with a Mass of	Contains
hydrogen	2.0	2.0 g H_2	6.02×10^{23} H_2 molecules or 1 mol of H_2 molecules (contains $2 \times 6.02 \times 10^{23}$ H atoms or 2 mol of H atoms)
oxygen	32.0	32.0 g O_2	6.02×10^{23} O_2 molecules or 1 mol of O_2 molecules (contains $2 \times 6.02 \times 10^{23}$ O atoms or 2 mol of O atoms)
methane	16.0	16.0 g CH_4	6.02×10^{23} CH_4 molecules or 1 mol of CH_4 molecules (contains 6.02×10^{23} C atoms and $4 \times 6.02 \times 10^{23}$ H atoms)
acetic acid (vinegar)	60.0	60.0 g CH_3COOH	6.02×10^{23} CH_3COOH molecules or 1 mol of CH_3COOH molecules

The amount of substance that contains the mass in grams numerically equal to its formula weight in amu contains 6.022×10^{23} formula units, or *one mole* of the substance. This is sometimes called the **molar mass** of the substance. Molar mass is *numerically equal* to the formula weight of the substance (the atomic weight for atoms of elements) and has the units grams/mole.

One mole of sodium hydroxide is 40.0 g of NaOH, and one mole of acetic acid is 60.0 g of CH_3COOH. One mole of any molecular substance contains 6.02×10^{23} molecules of the substance, as Table 2-5 illustrates.

Because no simple NaCl molecules exist at ordinary temperatures, it is inappropriate to refer to the "molecular weight" of NaCl or any ionic compound. One mole of an ionic compound contains 6.02×10^{23} *formula units* of the substance. Recall that one formula unit of sodium chloride consists of one sodium ion, Na^+, and one chloride ion, Cl^-. One mole, or 58.4 g, of NaCl contains 6.02×10^{23} Na^+ ions and 6.02×10^{23} Cl^- ions (Table 2-6).

The mole concept, together with Avogadro's number, provides important connections among the extensive properties mass of substance, number of moles of substance, and number of molecules or ions. These are summarized as follows.

Mass A $\longleftrightarrow$ **Moles A**

$$\text{FW A (molar mass A)} = \frac{\text{g A}}{\text{mol A}}$$

Moles A $\longleftrightarrow$ Formula units of A

$$\text{Avogadro's number} = \frac{\text{Formula units of A}}{\text{mol A}}$$

TABLE 2-6 *One Mole of Some Ionic Compounds*

Compound	Formula Weight	A Sample with a Mass of 1 Mol	Contains
sodium chloride	58.4	58.4 g NaCl	6.02×10^{23} Na^+ ions or 1 mol of Na^+ ions; 6.02×10^{23} Cl^- ions or 1 mol of Cl^- ions
calcium chloride	111.0	111.0 g $CaCl_2$	6.02×10^{23} Ca^{2+} ions or 1 mol of Ca^{2+} ions; $2(6.02 \times 10^{23})$ Cl^- ions or 2 mol of Cl^- ions
aluminum sulfate	342.1	342.1 g $Al_2(SO_4)_3$	$2(6.02 \times 10^{23})$ Al^{3+} ions or 2 mol of Al^{3+} ions; $3(6.02 \times 10^{23})$ SO_4^{2-} ions or 3 mol of SO_4^{2-} ions

The following examples show the relations between numbers of molecules, atoms, or formula units and their masses.

EXAMPLE 2-8 *Masses of Molecules*

What is the mass in grams of 10.0 million SO_2 molecules?

Plan

One mole of SO_2 contains 6.02×10^{23} SO_2 molecules and has a mass of 64.1 grams.

Solution

$$\underline{?}\text{ g } SO_2 = 10.0 \times 10^6 \text{ } SO_2 \text{ molecules} \times \frac{64.1 \text{ g } SO_2}{6.02 \times 10^{23} \text{ } SO_2 \text{ molecules}}$$

$$= \boxed{1.06 \times 10^{-15} \text{ g } SO_2}$$

When fewer than four significant figures are needed in calculations, Avogadro's number may be rounded off to 6.02×10^{23}.

Ten million SO_2 molecules have a mass of only 0.00000000000000106 g. Commonly used analytical balances are capable of weighing to ±0.0001 g.

You should now work Exercise 44.

EXAMPLE 2-9 *Moles*

How many (a) moles of O_2, (b) O_2 molecules, and (c) O atoms are contained in 40.0 g of oxygen gas (dioxygen) at 25°C?

Plan

We construct the needed unit factors from the following equalities: (a) the mass of one mole of O_2 is 32.0 g (molar mass O_2 = 32.0 g/mol); (b) one mole of O_2 contains 6.02×10^{23} O_2 molecules; (c) one O_2 molecule contains two O atoms.

Solution

One mole of O_2 contains 6.02×10^{23} O_2 molecules, and its mass is 32.0 g.

(a)
$$\underline{?} \text{ mol } O_2 = 40.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.0 \text{ g } O_2} = \boxed{1.25 \text{ mol } O_2}$$

(b)
$$\underline{?} \; O_2 \text{ molecules} = 40.0 \text{ g } O_2 \times \frac{6.02 \times 10^{23} \; O_2 \text{ molecules}}{32.0 \text{ g } O_2}$$
$$= \boxed{7.52 \times 10^{23} \text{ molecules}}$$

Or, we can use the number of moles of O_2 calculated in part (a) to find the number of O_2 molecules.

$$\underline{?} \; O_2 \text{ molecules} = 1.25 \text{ mol } O_2 \times \frac{6.02 \times 10^{23} \; O_2 \text{ molecules}}{1 \text{ mol } O_2} = \boxed{7.52 \times 10^{23} \; O_2 \text{ molecules}}$$

(c)
$$\underline{?} \text{ O atoms} = 40.0 \text{ g } O_2 \times \frac{6.02 \times 10^{23} \; O_2 \text{ molecules}}{32.0 \text{ g } O_2} \times \frac{2 \text{ O atoms}}{1 \; O_2 \text{ molecule}}$$
$$= \boxed{1.50 \times 10^{24} \text{ O atoms}}$$

You should now work Exercise 30.

EXAMPLE 2-10 *Numbers of Atoms*

Calculate the number of hydrogen atoms in 39.6 g of ammonium sulfate, $(NH_4)_2SO_4$.

Plan

In Example 2-10, we relate (a) grams to moles, (b) moles to formula units, and (c) formula units to H atoms.

One mole of $(NH_4)_2SO_4$ is 6.02×10^{23} formula units and has a mass of 132.1 g.

$$\text{g of } (NH_4)_2SO_4 \longrightarrow \text{mol of } (NH_4)_2SO_4 \longrightarrow \text{formula units of } (NH_4)_2SO_4 \longrightarrow \text{H atoms}$$

Solution

$$\underline{?} \text{ H atoms} = 39.6 \text{ g } (NH_4)_2SO_4 \times \frac{1 \text{ mol } (NH_4)_2SO_4}{132.1 \text{ g } (NH_4)_2SO_4} \times$$
$$\frac{6.02 \times 10^{23} \text{ formula units } (NH_4)_2SO_4}{1 \text{ mol } (NH_4)_2SO_4} \times \frac{8 \text{ H atoms}}{1 \text{ formula units } (NH_4)_2SO_4}$$
$$= \boxed{1.44 \times 10^{24} \text{ H atoms}}$$

You should now work Exercise 38.

2-8 PERCENT COMPOSITION AND FORMULAS OF COMPOUNDS

If the formula of a compound is known, its chemical composition can be expressed as the mass percent of each element in the compound (percent composition). For example, one carbon dioxide molecule, CO_2, contains one C atom and two O atoms. Percentage is the part divided by the whole times 100 percent (or simply parts per 100), so we can represent the percent composition of carbon dioxide as follows:

$$\%\ C = \frac{\text{mass of C}}{\text{mass of } CO_2} \times 100\% = \frac{\text{AW of C}}{\text{MW of } CO_2} \times 100\% = \frac{12.0 \text{ amu}}{44.0 \text{ amu}} \times 100\% = \boxed{27.3\%}$$

$$\%\ O = \frac{\text{mass of O}}{\text{mass of } CO_2} \times 100\% = \frac{2 \times \text{AW of O}}{\text{MW of } CO_2} \times 100\% = \frac{2(16.0 \text{ amu})}{44.0 \text{ amu}} \times 100\% = \boxed{72.7\%\ O}$$

One *mole* of CO_2 (44.0 g) contains one *mole* of C atoms (12.0 g) and two *moles* of O atoms (32.0 g). We could therefore have used these masses in the preceding calculation. These numbers are the same as the ones used—only the units are different. In Example 2-11 we shall base our calculation on one *mole* rather than one *molecule*.

As a check, we see that the percentages add to 100%.

Example 2-11 *Percent Composition*

Calculate the percent composition by mass of HNO_3.

Plan

We first calculate the mass of one mole as in Example 2-7. Then we express the mass of each element as a percent of the total.

Solution

The molar mass of HNO_3 is calculated first.

Number of Mol of Atoms		× Mass of One Mol of Atoms	= Mass Due to Element
1 × H =	1	× 1.0 g	= 1.0 g of H
1 × N =	1	× 14.0 g	= 14.0 g of N
3 × O =	3	× 16.0 g	= 48.0 g of O
		Mass of 1 mol of HNO_3	= 63.0 g

Now, its percent composition is

$$\%\ H = \frac{\text{mass of H}}{\text{mass of } HNO_3} \times 100\% = \frac{1.0 \text{ g}}{63.0 \text{ g}} \times 100\% = \boxed{1.6\%\ H}$$

$$\%\ N = \frac{\text{mass of N}}{\text{mass of } HNO_3} \times 100\% = \frac{14.0 \text{ g}}{63.0 \text{ g}} \times 100\% = \boxed{22.2\%\ N}$$

$$\%\ O = \frac{\text{mass of O}}{\text{mass of } HNO_3} \times 100\% = \frac{48.0 \text{ g}}{63.0 \text{ g}} \times 100\% = \boxed{76.2\%\ O}$$

$$\text{Total} = 100.0\%$$

When chemists use the % notation, they mean percent by mass unless they specify otherwise.

You should now work Exercise 46.

Nitric acid is 1.6% H, 22.2% N, and 76.2% O by mass. All samples of pure HNO_3 have this composition, according to the Law of Definite Proportions.

CHEMISTRY IN USE *The Development of Science*

Names of the Elements

If you were to discover a new element, how would you name it? Throughout history, scientists have answered this question in different ways. Most have chosen to honor a person or place or to describe the new substance.

Until the Middle Ages only nine elements were known: gold, silver, tin, mercury, copper, lead, iron, sulfur, and carbon. The metals' chemical symbols are taken from descriptive Latin names: *aurum* ("yellow"), *argentum* ("shining"), *stannum* ("dripping" or "easily melted"), *hydrargyrum* ("silvery water"), *cuprum* ("Cyprus," where many copper mines were located), *plumbum* (exact meaning unknown—possibly "heavy"), and *ferrum* (also unknown). Mercury is named after the planet, one reminder that the ancients associated metals with gods and celestial bodies. In turn, both the planet, which moves rapidly across the sky, and the element, which is the only metal that is liquid at room temperature and thus flows rapidly, are named for the fleet god of messengers in Roman mythology. In English, mercury is nicknamed "quicksilver."

Prior to the reforms of Antoine Lavoisier (1743–1794), chemistry was a largely nonquantitative, unsystematic science in which experimenters had little contact with each other. In 1787, Lavoisier published his *Methode de Nomenclature Chimique*, which proposed, among other changes, that all new elements be named descriptively. For the next 125 years, most elements were given names that corresponded to their properties. Greek roots were one popular source, as evidenced by hydrogen (*hydros-gen*, "water-producing"), oxygen (*oksys-gen*, "acid-producing"), nitrogen (*nitron-gen*, "soda-producing"), bromine (*bromos*, "stink"), and argon (*a-er-gon*, "no reaction"). The discoverers of argon, Sir William Ramsay (1852–1916) and Baron Rayleigh (1842–1919), originally proposed the name *aeron* (from *aer* or air), but critics thought it was too close to the biblical name Aaron! Latin roots, such as *radius* ("ray"), were also used (radium and radon are both naturally radioactive elements that emit "rays"). Color was often the determining property, especially after the invention of the spectroscope in 1859, because different elements (or the light that they emit) have prominent characteristic colors. Cesium, indium, iodine, rubidium, and thallium were all named in this manner. Their respective Greek and Latin roots denote blue-gray, indigo, violet, red, and green (*thallus* means "tree sprout"). Because of the great variety of colors of its compounds, iridium takes its name from the Latin *iris*, meaning "rainbow." Alternatively, an element name might suggest a mineral or the ore that contained it. One example is Wolfram or tungsten (W), which was isolated from wolframite. Two other "inconsistent" elemental symbols, K and Na, arose from occurrence as well. *Kalium* was first obtained from the saltwort plant, *Salsola kali*, and *natrium* from niter. Their English names, potassium and sodium, are derived from the ores potash and soda.

Other elements, contrary to Lavoisier's suggestion, were named after planets, mythological figures, places, or superstitions. "Celestial elements" include helium ("sun"), tellurium ("earth"), selenium ("moon"—the element was discovered in close proximity to tellurium), cerium (the asteroid Ceres, which was discovered only two years before the element), and uranium (the planet Uranus, discovered a few years earlier). The first two transuranium elements (those *beyond* uranium) to be produced were named neptunium and plutonium for the next two planets, Neptune and Pluto. The names promethium (Prometheus, who stole fire from heaven), vanadium (Scandinavian goddess Vanadis), titanium (Titans, the first sons of the earth), tantalum (Tantalos, father of the Greek goddess Niobe), and thorium (Thor, Scandinavian god of war) all arise from Greek or Norse mythology.

"Geographical elements," shown on the map, sometimes honored the discoverer's native country or workplace. The Latin names for Russia (*ruthenium*), France (*gallium*), Paris (*lutetium*), and Germany (*germanium*) were among those used. Marie Sklodowska Curie named one of the elements that she discovered polonium, after her native Poland. Often the locale of discovery lends its name to the element; the record holder is certainly the Swedish village Ytterby, the site of ores from which the four elements terbium, erbium, ytterbium, and yttrium were isolated. Elements honoring important scientists include curium, einsteinium, nobelium, fermium, and lawrencium.

Most of the elements now known were given titles peacefully, but a few were not. Niobium, isolated in 1803 by Ekeberg from an ore that also contained tantalum, and named after Niobe (daughter of Tantalos), was later found to be identical to an 1802 discovery of C. Hatchett, columbium. (Interestingly, Hatchett first found the element in an ore sample that had been sent to England more than a century

Many chemical elements were named after places.

earlier by John Winthrop, the first governor of Connecticut.) Although "niobium" became the accepted designation in Europe, the Americans, not surprisingly, chose "columbium." It was not until 1949—when the International Union of Pure and Applied Chemistry (IUPAC) ended more than a century of controversy by ruling in favor of mythology—that element 41 received a unique name.

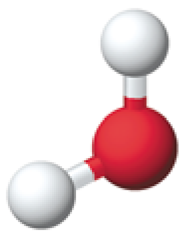
In 1978, the IUPAC recommended that elements beyond 103 be known temporarily by systematic names based on numerical roots; element 104 is unnilquadium (*un* for 1, *nil* for 0, *quad* for 4, plus the *-ium* ending), followed by unnilpentium, unnilhexium, and so on. Arguments over the names of elements 104 and 105 prompted the IUPAC to begin hearing claims of priority to numbers 104 to 109. The IUPAC's final recommendations for these element names were announced in 1997. The names and symbols recommended by that report are: element 104, rutherfordium, Rf; element 105, dubnium, Db; element 106, seaborgium, Sg; element 107, bohrium, Bh; element 108, hassium, Hs; and element 109, meitnerium, Mt. Some of these (Rf and Bh) are derived from the names of scientists prominent in the development of atomic theory; others (Sg, Hs, and Mt) are named for scientists who were involved in the discovery of heavy elements. Dubnium is named in honor of the Dubna laboratory in the former Soviet Union, where important contributions to the creation of heavy elements have originated.

Lisa Saunders Boffa
Senior Chemist
Exxon Corporation

Problem-Solving Tip: ***The Whole Is Equal to the Sum of Its Parts***

Percentages must add to 100%. Roundoff errors may not cancel, however, and totals such as 99.9% or 100.1% may be obtained in calculations. As an alternative method of calculation, if we know all of the percentages except one, we can subtract their sum from 100% to obtain the other value.

2-9 DERIVATION OF FORMULAS FROM ELEMENTAL COMPOSITION

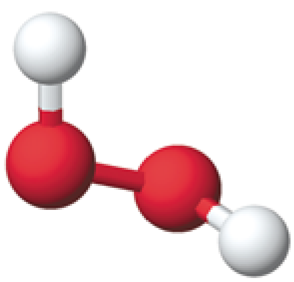


A ball-and-stick model of a molecule of water, H_2O.

The **simplest,** or **empirical, formula** for a compound is the smallest whole-number ratio of atoms present. For molecular compounds the **molecular formula** indicates the *actual* numbers of atoms present in a molecule of the compound. It may be the same as the simplest formula or else some whole-number multiple of it. For example, the simplest and molecular formulas for water are both H_2O; however, for hydrogen peroxide, they are HO and H_2O_2, respectively.

Each year thousands of new compounds are made in laboratories or discovered in nature. One of the first steps in characterizing a new compound is the determination of its percent composition. A *qualitative* analysis is performed to determine *which* elements are present in the compound. Then a *quantitative* analysis is performed to determine the *amount* of each element.

Once the percent composition of a compound (or its elemental composition by mass) is known, the simplest formula can be determined.



A ball-and-stick model of a molecule of hydrogen peroxide, H_2O_2.

Example 2-12 *Simplest Formulas*

Compounds containing sulfur and oxygen are serious air pollutants; they represent the major cause of acid rain. Analysis of a sample of a pure compound reveals that it contains 50.1% sulfur and 49.9% oxygen by mass. What is the simplest formula of the compound?

Plan

One mole of atoms of any element is 6.02×10^{23} atoms, so the ratio of moles of atoms in any sample of a compound is the same as the ratio of atoms in that compound. This calculation is carried out in two steps.

Remember that *percent* means *parts per hundred.*

Step 1: Let's consider 100.0 g of compound, which contains 50.1 g of S and 49.9 g of O. We calculate the number of moles of atoms of each.

Step 2: We then obtain a whole-number ratio between these numbers that gives the ratio of atoms in the sample and hence in the simplest formula for the compound.

Solution

Step 1:

$$\underline{?}\text{ mol S atoms} = 50.1\text{ g S} \times \frac{1\text{ mol S atoms}}{32.1\text{ g S}} = 1.56\text{ mol S atoms}$$

$$\underline{?}\text{ mol O atoms} = 49.9\text{ g O} \times \frac{1\text{ mol O atoms}}{16.0\text{ g O}} = 3.12\text{ mol O atoms}$$

Step 2: Now we know that 100.0 g of the compound contains 1.56 mol of S atoms and 3.12 mol of O atoms. We obtain a whole-number ratio between these numbers that gives the ratio of atoms in the simplest formula.

$$\frac{1.56}{1.56} = 1.00 \text{ S}$$

$$\frac{3.12}{1.56} = 2.00 \text{ O}$$

$$SO_2$$

You should now work Exercise 52.

A simple and useful way to obtain whole-number ratios among several numbers follows: (a) Divide each number by the smallest, and then, (b) if necessary, multiply all of the resulting numbers by the smallest whole number that will eliminate fractions.

The solution for Example 2-12 can be set up in tabular form.

Element	Relative Mass of Element	Relative Number of Atoms (divide mass by AW)	Divide by Smaller Number	Smallest Whole-Number Ratio of Atoms
S	50.1	$\frac{50.1}{32.1} = 1.56$	$\frac{1.56}{1.56} = 1.00$ S	SO_2
O	49.9	$\frac{49.9}{16.0} = 3.12$	$\frac{3.12}{1.56} = 2.00$ O	

The "Relative Mass" column is proportional to the mass of each element in grams. With this interpretation, the next column could be headed "Relative Number of *Moles* of Atoms." Then the last column would represent the smallest whole-number ratios of *moles* of atoms. But because a mole is always the same number of items (atoms), that ratio is the same as the smallest whole-number ratio of atoms.

This tabular format provides a convenient way to solve simplest-formula problems, as the next example illustrates.

EXAMPLE 2-13 *Simplest Formula*

A 20.882-gram sample of an ionic compound is found to contain 6.072 grams of Na, 8.474 grams of S, and 6.336 grams of O. What is its simplest formula?

Plan

We reason as in Example 2-12, calculating the number of moles of each element and the ratio among them. Here we use the tabular format that was introduced earlier.

Solution

Element	Relative Mass of Element	Relative Number of Atoms (divide mass by AW)	Divide by Smallest Number	Convert Fractions to Whole Numbers (multiply by integer)	Smallest Whole-Number Ratio of Atoms
Na	6.072	$\frac{6.072}{23.0} = 0.264$	$\frac{0.264}{0.264} = 1.00$	$1.00 \times 2 = 2$ Na	
S	8.474	$\frac{8.474}{32.1} = 0.264$	$\frac{0.264}{0.264} = 1.00$	$1.00 \times 2 = 2$ S	$Na_2S_2O_3$
O	6.336	$\frac{6.336}{16.0} = 0.396$	$\frac{0.396}{0.264} = 1.50$	$1.50 \times 2 = 3$ O	

The ratio of atoms in the simplest formula *must be a whole-number ratio* (by definition). To convert the ratio 1 : 1 : 1.5 to a whole-number ratio, each number in the ratio was multiplied by 2, which gave the simplest formula $Na_2S_2O_3$.

You should now work Exercise 53.

Problem-Solving Tip: *Know Common Fractions in Decimal Form*

As Example 2-13 illustrates, sometimes we must convert a fraction to a whole number by multiplication by the correct integer. But we must first recognize which fraction is represented by a nonzero part of a number. The decimal equivalents of the following fractions may be useful.

Decimal Equivalent (to 2 places)	Fraction	To Convert to Integer, Multiply by
0.50	$\frac{1}{2}$	2
0.33	$\frac{1}{3}$	3
0.67	$\frac{2}{3}$	3
0.25	$\frac{1}{4}$	4
0.75	$\frac{3}{4}$	4
0.20	$\frac{1}{5}$	5

The fractions $\frac{2}{5}$, $\frac{3}{5}$, $\frac{4}{5}$ are equal to 0.40, 0.60, and 0.80, respectively; these should be multiplied by 5.

When we use the procedure given in this section, we often obtain numbers such as 0.99 and 1.52. Because the results obtained by analysis of samples usually contain some error (as well as roundoff errors), we would interpret 0.99 as 1.0 and 1.52 as 1.5.

Millions of compounds are composed of carbon, hydrogen, and oxygen. Analyses for C and H can be performed in a C-H combustion system (Figure 2-10). An accurately

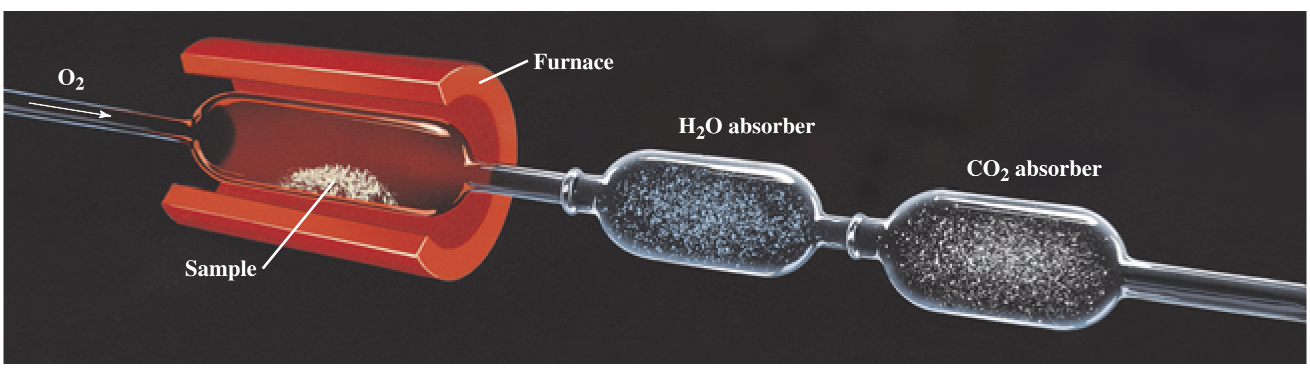


Figure 2-10 A combustion train used for carbon-hydrogen analysis. The absorbent for water is magnesium perchlorate, $Mg(ClO_4)_2$. Carbon dioxide is absorbed by finely divided sodium hydroxide supported on glass wool. Only a few milligrams of sample is needed for analysis.

known mass of a compound is burned in a furnace in a stream of oxygen. The carbon and hydrogen in the sample are converted to carbon dioxide and water vapor, respectively. The resulting increases in masses of the CO_2 and H_2O absorbers can then be related to the masses and percentages of carbon and hydrogen in the original sample.

EXAMPLE 2-14 *Percent Composition*

Hydrocarbons are organic compounds composed entirely of hydrogen and carbon. A 0.1647-gram sample of a pure hydrocarbon was burned in a C-H combustion train to produce 0.4931 gram of CO_2 and 0.2691 gram of H_2O. Determine the masses of C and H in the sample and the percentages of these elements in this hydrocarbon.

Hydrocarbons are obtained from coal and coal tar and from oil and gas wells. The main use of hydrocarbons is as fuels. The simplest hydrocarbons are

methane	CH_4
ethane	C_2H_6
propane	C_3H_8
butane	C_4H_{10}

Plan

Step 1: We use the observed mass of CO_2, 0.4931 grams, to determine the mass of carbon in the original sample. There is one mole of carbon atoms, 12.01 grams, in each mole of CO_2, 44.01 grams; we use this information to construct the unit factor

$$\frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2}$$

Step 2: Likewise, we can use the observed mass of H_2O, 0.2691 grams, to calculate the amount of hydrogen in the original sample. We use the fact that there are two moles of hydrogen atoms, 2.016 grams, in each mole of H_2O, 18.02 grams, to construct the unit factor

$$\frac{2.016 \text{ g H}}{18.02 \text{ g } H_2O}$$

We could calculate the mass of H by subtracting mass of C from mass of sample. It is good experimental practice, however, when possible, to base both on experimental measurements, as we have done here. This would help to check for errors in the analysis or calculation.

Step 3: Then we calculate the percentages by mass of each element in turn, using the relationship

$$\% \text{ element} = \frac{\text{g element}}{\text{g sample}} \times 100\%$$

Solution

Step 1: $\underline{?}$ g C $= 0.4931 \text{ g } CO_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} = \boxed{0.1346 \text{ g C}}$

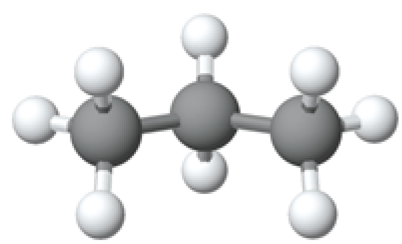
Step 2: $\underline{?}$ g H $= 0.2691 \text{ g } H_2O \times \frac{2.016 \text{ g H}}{18.02 \text{ g } H_2O} = \boxed{0.03010 \text{ g H}}$

Step 3: $\% \text{ C} = \frac{0.1346 \text{ g C}}{0.1647 \text{ g sample}} \times 100\% = \boxed{81.72\% \text{ C}}$

$\% \text{H} = \frac{0.03010 \text{ g H}}{0.1647 \text{ g sample}} \times 100\% = \boxed{18.28\% \text{ H}}$

Total = 100.00%

You should now work Exercise 58.



Can you show that the hydrocarbon in Example 2-14 is propane, C_3H_8?

When the compound to be analyzed contains oxygen, the calculation of the amount or percentage of oxygen in the sample is somewhat different. Part of the oxygen that goes to form CO_2 and H_2O comes from the sample and part comes from the oxygen stream supplied. For that reason, we cannot directly determine the amount of oxygen already in the sample. The approach is to analyze as we did in Example 2-14 for all elements *except* oxygen. Then we subtract the sum of their masses from the mass of the original sample to obtain the mass of oxygen. The next example illustrates such a calculation.

EXAMPLE 2-15 *Percent Composition*

Glucose, a simple sugar, is the main component of intravenous feeding liquids. Its common name is dextrose. It is also one of the products of carbohydrate metabolism.

We say that the mass of O in the sample is calculated by *difference.*

A 0.1014-g sample of purified glucose was burned in a C-H combustion train to produce 0.1486 g of CO_2 and 0.0609 g of H_2O. An elemental analysis showed that glucose contains only carbon, hydrogen, and oxygen. Determine the masses of C, H, and O in the sample and the percentages of these elements in glucose.

Plan

Steps 1 and 2: We first calculate the masses of carbon and hydrogen as we did in Example 2-14.
Step 3: The rest of the sample must be oxygen because glucose has been shown to contain only C, H, and O. So we subtract the masses of C and H from the total mass of sample.
Step 4: Then we calculate the percentage by mass for each element.

Solution

Step 1: $\underline{?}\text{ g C} = 0.1486\text{ g CO}_2 \times \dfrac{12.01\text{ g C}}{44.01\text{ g CO}_2} = \boxed{0.04055\text{ g C}}$

Step 2: $\underline{?}\text{ g H} = 0.0609\text{ g H}_2\text{O} \times \dfrac{2.016\text{ g H}}{18.02\text{ g H}_2\text{O}} = \boxed{0.00681\text{ g H}}$

Step 3: $\underline{?}\text{ g O} = 0.1014\text{ g sample} - [0.04055\text{ g C} + 0.00681\text{ g H}] = \boxed{0.0540\text{ g O}}$

Step 4: Now we can calculate the percentages by mass for each element:

$$\%\text{ C} = \frac{0.04055\text{ g C}}{0.1014\text{ g}} \times 100\% = \boxed{39.99\%\text{ C}}$$

$$\%\text{ H} = \frac{0.00681\text{ g H}}{0.1014\text{ g}} \times 100\% = \boxed{6.72\%\text{ H}}$$

$$\%\text{ O} = \frac{0.0540\text{ g O}}{0.1014\text{ g}} \times 100\% = \boxed{53.2\%\text{ O}}$$

$$\text{Total} = 99.9\%$$

You should now work Exercise 60.

2-10 DETERMINATION OF MOLECULAR FORMULAS

Percent composition data yield only simplest formulas. To determine the molecular formula for a molecular compound, *both* its simplest formula and its molecular weight must be known. Some methods for experimental determination of molecular weights are introduced in Chapters 12 and 14.

For many compounds the molecular formula is a multiple of the simplest formula. Consider butane, C_4H_{10}. The simplest formula for butane is C_2H_5, but the molecular formula contains twice as many atoms; that is, $2 \times (C_2H_5) = C_4H_{10}$. Benzene, C_6H_6, is another example. The simplest formula for benzene is CH, but the molecular formula contains six times as many atoms; that is, $6 \times (CH) = C_6H_6$.

The molecular formula for a compound is either the same as, or an *integer* multiple of, the simplest formula.

$$\text{Molecular formula} = n \times \text{simplest formula}$$

So we can write

$$\text{Molecular weight} = n \times \text{simplest formula weight}$$

$$n = \frac{\text{molecular weight}}{\text{simplest formula weight}}$$

The molecular formula is then obtained by multiplying the simplest formula by the integer, n.

EXAMPLE 2-16 *Molecular Formula*

In Example 2-15, we found the elemental composition of glucose. Other experiments show that its molecular weight is approximately 180 amu. Determine the simplest formula and the molecular formula of glucose.

Plan

Step 1: We first use the masses of C, H, and O found in Example 2-15 to determine the simplest formula.

Step 2: We can use the simplest formula to calculate the simplest formula weight. Because the molecular weight of glucose is known (approximately 180 amu), we can determine the molecular formula by dividing the molecular weight by the simplest formula weight.

$$n = \frac{\text{molecular weight}}{\text{simplest formula weight}}$$

The molecular weight is n times the simplest formula weight, so the molecular formula of glucose is n times the simplest formula.

As an alternative, we could have used the percentages by mass from Example 2-15. Using the earliest available numbers helps to minimize the effects of rounding-off errors.

Solution

Step 1:

Element	Mass of Element	Moles of Element (divide mass by AW)	Divide by Smallest	Smallest Whole-Number Ratio of Atoms
C	0.04055 g	$\frac{0.04055}{12.01} = 0.003376$ mol	$\frac{0.003376}{0.003376} = 1.00$ C	
H	0.00681 g	$\frac{0.00681}{1.008} = 0.00676$ mol	$\frac{0.00676}{0.003376} = 2.00$ H	CH_2O
O	0.0540 g	$\frac{0.0540}{16.00} = 0.00338$ mol	$\frac{0.00338}{0.003376} = 1.00$ O	

Many sugars are rich sources in our diet. The most familiar is ordinary table sugar, which is sucrose, $C_{12}H_{22}O_{11}$. An enzyme in our saliva readily splits sucrose into two simple sugars, glucose and fructose. The simplest formula for both glucose and fructose is $C_6H_{12}O_6$. They have different structures and different properties, however, so they are different compounds.

Step 2: The simplest formula is CH_2O, which has a formula weight of 30.03 amu. Because the molecular weight of glucose is approximately 180 amu, we can determine the molecular formula by dividing the molecular weight by the simplest formula weight.

$$n = \frac{180\text{ amu}}{30.03\text{ amu}} = 6.00$$

The molecular weight is six times the simplest formula weight, $6 \times (CH_2O) = C_6H_{12}O_6$, so the molecular formula of glucose is $C_6H_{12}O_6$.

You should now work Exercises 63 and 64.

As we shall see when we discuss the composition of compounds in some detail, two (and sometimes more) elements may form more than one compound. The **Law of Multiple Proportions** summarizes many experiments on such compounds. It is usually stated: When two elements, A and B, form more than one compound, the ratio of the masses of element B that combine with a given mass of element A in each of the compounds can be expressed by small whole numbers. Water, H_2O, and hydrogen peroxide, H_2O_2, provide an example. The ratio of masses of oxygen that combine with a given mass of hydrogen in H_2O and H_2O_2 is 1 : 2. Many similar examples, such as CO and CO_2 (1 : 2 ratio) and SO_2 and SO_3 (2 : 3 ratio), are known. The Law of Multiple Proportions had been recognized from studies of elemental composition before the time of Dalton. It provided additional support for his atomic theory.

EXAMPLE 2-17 *Law of Multiple Proportions*

What is the ratio of the masses of oxygen that are combined with one gram of nitrogen in the compounds N_2O_3 and NO?

Plan

First we calculate the mass of O that combines with one gram of N in each compound. Then we determine the ratio of the values of $\frac{\text{g O}}{\text{g N}}$ for the two compounds.

Solution

In N_2O_3: $\frac{\underline{?}\text{ g O}}{\text{g N}} = \frac{48.0\text{ g O}}{28.0\text{ g N}} = 1.71\text{ g O/g N}$

In NO: $\frac{\underline{?}\text{ g O}}{\text{g N}} = \frac{16.0\text{ g O}}{14.0\text{ g N}} = 1.14\text{ g O/g N}$

$$\text{The ratio is } \frac{\frac{\text{g O}}{\text{g N}}\ (\text{in } N_2O_3)}{\frac{\text{g O}}{\text{g N}}\ (\text{in NO})} \rightarrow \frac{1.71\text{ g O/g N}}{1.14\text{ g O/g N}} = \frac{1.5}{1.0} = \frac{3}{2}$$

We see that the ratio is three mass units of O (in N_2O_3) to two mass units of O (in NO).

You should now work Exercises 67 and 68.

2-11 SOME OTHER INTERPRETATIONS OF CHEMICAL FORMULAS

Once we master the mole concept and the meaning of chemical formulas, we can use them in many other ways. The examples in this section illustrate a few additional kinds of information we can get from a chemical formula and the mole concept.

EXAMPLE 2-18 *Composition of Compounds*

What mass of chromium is contained in 35.8 g of $(NH_4)_2Cr_2O_7$?

Plan

Let us first solve the problem in several steps.

Step 1: The formula tells us that each mole of $(NH_4)_2Cr_2O_7$ contains two moles of Cr atoms, so we first find the number of moles of $(NH_4)_2Cr_2O_7$, using the unit factor

$$\frac{1 \text{ mol } (NH_4)_2Cr_2O_7}{252.0 \text{ g } (NH_4)_2Cr_2O_7}$$

Step 2: Then we convert the number of moles of $(NH_4)_2Cr_2O_7$ into the number of moles of Cr atoms it contains, using the unit factor

$$\frac{2 \text{ mol Cr atoms}}{1 \text{ mol } (NH_4)_2Cr_2O_7}$$

Step 3: We then use the atomic weight of Cr to convert the number of moles of chromium atoms to mass of chromium.

$$\text{Mass } (NH_4)_2Cr_2O_7 \longrightarrow \text{mol } (NH_4)_2Cr_2O_7 \longrightarrow \text{mol Cr} \longrightarrow \text{Mass Cr}$$

Solution

Step 1:
$$\underline{?} \text{ mol } (NH_4)_2Cr_2O_7 = 35.8 \text{ g } (NH_4)_2Cr_2O_7 \times \frac{1 \text{ mol } (NH_4)_2Cr_2O_7}{252.0 \text{ g } (NH_4)_2Cr_2O_7} = \boxed{0.142 \text{ mol } (NH_4)_2Cr_2O_7}$$

Step 2:
$$\underline{?} \text{ mol Cr atoms} = 0.142 \text{ mol } (NH_4)_2Cr_2O_7 \times \frac{2 \text{ mol Cr atoms}}{1 \text{ mol } (NH_4)_2Cr_2O_7} = \boxed{0.284 \text{ mol Cr atoms}}$$

Step 3:
$$\underline{?} \text{ g Cr} = 0.284 \text{ mol Cr atoms} \times \frac{52.0 \text{ g Cr}}{1 \text{ mol Cr atoms}} = \boxed{14.8 \text{ g Cr}}$$

If you understand the reasoning in these conversions, you should be able to solve this problem in a single setup:

$$\underline{?} \text{ g Cr} = 35.8 \text{ g } (NH_4)_2Cr_2O_7 \times \frac{1 \text{ mol } (NH_4)_2Cr_2O_7}{252.0 \text{ g } (NH_4)_2Cr_2O_7} \times \frac{2 \text{ mol Cr atoms}}{1 \text{ mol } (NH_4)_2Cr_2O_7} \times \frac{52.0 \text{ g Cr}}{1 \text{ mol Cr}} = \boxed{14.8 \text{ g Cr}}$$

You should now work Exercise 72.

EXAMPLE 2-19 *Composition of Compounds*

What mass of potassium chlorate, $KClO_3$, would contain 40.0 grams of oxygen?

Plan

The formula $KClO_3$ tells us that each mole of $KClO_3$ contains three moles of oxygen atoms. Each mole of oxygen atoms weighs 16.0 grams. So we can set up the solution to convert:

$$\text{Mass O} \longrightarrow \text{mol O} \longrightarrow \text{mol } KClO_3 \longrightarrow \text{Mass } KClO_3$$

Solution

$$\underline{?}\text{ g } KClO_3 = 40.0\text{ g O} \times \frac{1\text{ mol O atoms}}{16.0\text{ g O atoms}} \times \frac{1\text{ mol } KClO_3}{3\text{ mol O atoms}} \times \frac{122.6\text{ g } KClO_3}{1\text{ mol } KClO_3}$$

$$= \boxed{102\text{ g } KClO_3}$$

You should now work Exercise 74.

✓ Problem-Solving Tip: *How Do We Know When . . . ?*

How do we know when to represent oxygen as O and when as O_2? A *compound* that contains oxygen *does not* contain O_2 molecules. So we solve problems such as Example 2-19 in terms of moles of O atoms. Thus, we must use the formula weight for O, which is 16.0 g O atoms/1 mol O atoms. Similar reasoning applies to compounds containing other elements that are polyatomic molecules in *pure elemental form*, such as H_2, Cl_2, or P_4.

EXAMPLE 2-20 *Composition of Compounds*

(a) What mass of sulfur dioxide, SO_2, would contain the same mass of oxygen as is contained in 33.7 g of arsenic pentoxide, As_2O_5?
(b) What mass of calcium chloride, $CaCl_2$, would contain the same number of chloride ions as are contained in 48.6 g of sodium chloride, NaCl?

Plan

(a) We could find explicitly the number of grams of O in 33.7 g of As_2O_5, and then find the mass of SO_2 that contains that same number of grams of O. But this method includes some unnecessary calculation. We need only convert to *moles* of O (because this is the same amount of O regardless of its environment) and then to moles of SO_2 to obtain mass of SO_2.

$$\text{Mass } As_2O_5 \longrightarrow \text{mol } As_2O_5 \longrightarrow \text{mol O atoms} \longrightarrow \text{mol } SO_2 \longrightarrow \text{Mass } SO_2$$

(b) Because one mole always consists of the same number (Avogadro's number) of items, we can reason in terms of *moles* of Cl^- ions and solve as in part (a).

$$\text{Mass NaCl} \longrightarrow \text{mol NaCl} \longrightarrow \text{mol } Cl^- \text{ ions} \longrightarrow \text{mol } CaCl_2 \longrightarrow \text{Mass } CaCl_2$$

Solution

(a)

$$\underline{?}\text{ g } SO_2 = 33.7\text{ g } As_2O_5 \times \frac{1\text{ mol } As_2O_5}{229.8\text{ g } As_2O_5} \times \frac{5\text{ mol O atoms}}{1\text{ mol } As_2O_5}$$

$$\times \frac{1\text{ mol } SO_2}{2\text{ mol O atoms}} \times \frac{64.1\text{ g } SO_2}{1\text{ mol } SO_2} = \boxed{23.5\ SO_2}$$

(b)

$$\underline{\ ?\ }\text{ g CaCl}_2 = 48.6\text{ NaCl} \times \frac{1\text{ mol NaCl}}{58.4\text{ g NaCl}} \times \frac{1\text{ mol Cl}^-}{1\text{ mol NaCl}}$$

$$\times \frac{1\text{ mol CaCl}_2}{2\text{ mol Cl}^-} \times \frac{111.0\text{ g CaCl}_2}{1\text{ mol CaCl}_2} = \boxed{46.2\text{ g CaCl}_2}$$

You should now work Exercise 76.

Figure 2-11 One mole of some compounds. The colorless liquid is water, H_2O (1 mol = 18.0 g = 18.0 mL). The white solid *(front left)* is *anhydrous* oxalic acid, $(COOH)_2$ (1 mol = 90.0 g). The second white solid *(front right)* is *hydrated* oxalic acid, $(COOH)_2 \cdot 2H_2O$ (1 mol = 126.0 g). The blue solid is hydrated copper(II) sulfate, $CuSO_4 \cdot 5H_2O$ (1 mol = 249.7 g). The red solid is mercury(II) oxide (1 mol = 216.6 g).

The physical appearance of one mole of each of some compounds is illustrated in Figure 2-11. Two different forms of oxalic acid are shown. The formula unit (molecule) of oxalic acid is $(COOH)_2$ (FW = 90.0 amu; molar mass = 90.0 g/mol). When oxalic acid is obtained by crystallization from a water solution, however, two molecules of water are present for each molecule of oxalic acid, even though it appears dry. The formula of this **hydrate** is $(COOH)_2 \cdot 2H_2O$ (FW = 126.1 amu; molar mass = 126.1 g/mol). The dot shows that the crystals contain two H_2O molecules per $(COOH)_2$ molecule. The water can be driven out of the crystals by heating to leave **anhydrous** oxalic acid, $(COOH)_2$. Anhydrous means "without water." Copper(II) sulfate, an *ionic* compound, shows similar behavior. Anhydrous copper(II) sulfate ($CuSO_4$; FW = 159.6 amu; molar mass = 159.6 g/mol) is almost white. Hydrated copper(II) sulfate ($CuSO_4 \cdot 5H_2O$; FW = 249.7 amu; molar mass = 249.7 g/mol) is deep blue. The following example illustrates how we might find and use the formula of a hydrate.

Heating blue $CuSO_4 \cdot 5H_2O$ forms anhydrous $CuSO_4$, which is gray. Some blue $CuSO_4 \cdot 5H_2O$ is visible in the cooler center portion of the crucible.

EXAMPLE 2-21 *Composition of Compounds*

A reaction requires pure anhydrous calcium sulfate, $CaSO_4$. Only an unidentified hydrate of calcium sulfate, $CaSO_4 \cdot xH_2O$, is available.

(a) We heat 67.5 g of unknown hydrate until all the water has been driven off. The resulting mass of pure $CaSO_4$ is 53.4 g. What is the formula of the hydrate, and what is its formula weight?

(b) Suppose we wish to obtain enough of this hydrate to supply 95.5 grams of $CaSO_4$. How many grams should we weigh out?

Plan

(a) To determine the formula of the hydrate, we must find the value of x in the formula $CaSO_4 \cdot xH_2O$. The mass of water removed from the sample is equal to the difference in the two masses given. The value of x is the number of moles of H_2O per mole of $CaSO_4$ in the hydrate.

(b) The formula weights of $CaSO_4$, 136.2 g/mol, and of $CaSO_4 \cdot xH_2O$, (136.2 + x18.0) g/mol, allow us to write the conversion factor required for the calculation.

Solution

(a) $\underline{?}$ g water driven off = 67.5 g $CaSO_4 \cdot xH_2O$ − 53.4 g $CaSO_4$ = 14.1 g H_2O

$$x = \frac{\underline{?}\text{ mol } H_2O}{\text{mol } CaSO_4} = \frac{14.1\text{ g } H_2O}{53.4\text{ g } CaSO_4} \times \frac{1\text{ mol } H_2O}{18.0\text{ g } H_2O} \times \frac{136.2\text{ g } CaSO_4}{1\text{ mol } CaSO_4} = \frac{2.00\text{ mol } H_2O}{\text{mol } CaSO_4}$$

Thus, the formula of the hydrate is $CaSO_4 \cdot 2H_2O$. Its formula weight is

$$\text{FW} = 1 \times (\text{formula weight } CaSO_4) + 2 \times (\text{formula weight } H_2O)$$
$$= 136.2\text{ g/mol} + 2(18.0\text{ g/mol}) = 172.2\text{ g/mol}$$

(b) The formula weights of $CaSO_4$ (136.2 g/mol) and of $CaSO_4 \cdot 2H_2O$ (172.2 g/mol) allow us to write the unit factor

$$\frac{172.2\text{ g } CaSO_4 \cdot 2H_2O}{136.2\text{ g } CaSO_4}$$

We use this factor to perform the required conversion:

$$\underline{?}\text{ g } CaSO_4 \cdot 2H_2O = 95.5\text{ g } CaSO_4\text{ desired} \times \frac{172.2\text{ g } CaSO_4 \cdot 2H_2O}{136.2\text{ g } CaSO_4}$$
$$= 121\text{ g } CaSO_4 \cdot 2H_2O$$

You should now work Exercise 78.

2-12 PURITY OF SAMPLES

Impurities are not necessarily bad. For example, inclusion of 0.02% KI, potassium iodide, in ordinary table salt has nearly eliminated goiter in the United States. Goiter is a disorder of the thyroid gland caused by a deficiency of iodine. Mineral water tastes better than purer, distilled water.

Most substances obtained from laboratory reagent shelves are not 100% pure. The **percent purity** is the mass percentage of a specified substance in an impure sample. When impure samples are used for precise work, account must be taken of impurities. The photo in the margin shows the label from reagent-grade sodium hydroxide, NaOH, which is 98.2% pure by mass. From this information we know that total impurities represent 1.8% of the mass of this material. We can write several unit factors:

$$\frac{98.2\text{ g NaOH}}{100\text{ g sample}}, \quad \frac{1.8\text{ g impurities}}{100\text{ g sample}}, \quad \text{and} \quad \frac{1.8\text{ g impurities}}{98.2\text{ g NaOH}}$$

The inverse of each of these gives us a total of six unit factors.

EXAMPLE 2-22 *Percent Purity*

Calculate the masses of NaOH and impurities in 45.2 g of 98.2% pure NaOH.

Plan

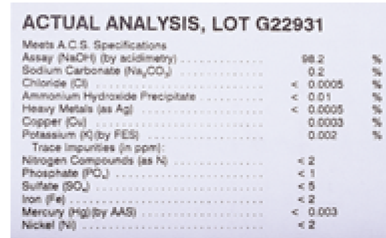
The percentage of NaOH in the sample gives the unit factor $\frac{98.2\text{ g NaOH}}{100\text{ g sample}}$. The remainder of the sample is 100% − 98.2% = 1.8% impurities; this gives the unit factor $\frac{1.8 \text{ g impurities}}{100 \text{ g sample}}$.

Solution

$$\underline{?} \text{ g NaOH} = 45.2 \text{ g sample} \times \frac{98.2 \text{ g NaOH}}{100 \text{ g sample}} = 44.4 \text{ g NaOH}$$

$$\underline{?} \text{ g impurities} = 45.2 \text{ g sample} \times \frac{1.8 \text{ g impurities}}{100 \text{ g sample}} = 0.81 \text{ g impurities}$$

You should now work Exercises 82 and 83.



A label from a bottle of sodium hydroxide, NaOH.

Problem-Solving Tip: *Utility of the Unit Factor Method*

Observe the beauty of the unit factor approach to problem solving! Such questions as "do we multiply by 0.982 or divide by 0.982?" never arise. The units always point toward the correct answer because we use unit factors constructed so that units *always* cancel out until we arrive at the desired unit.

Many important relationships have been introduced in this chapter. Some of the most important transformations you have seen in Chapters 1 and 2 are summarized in Figure 2-12.

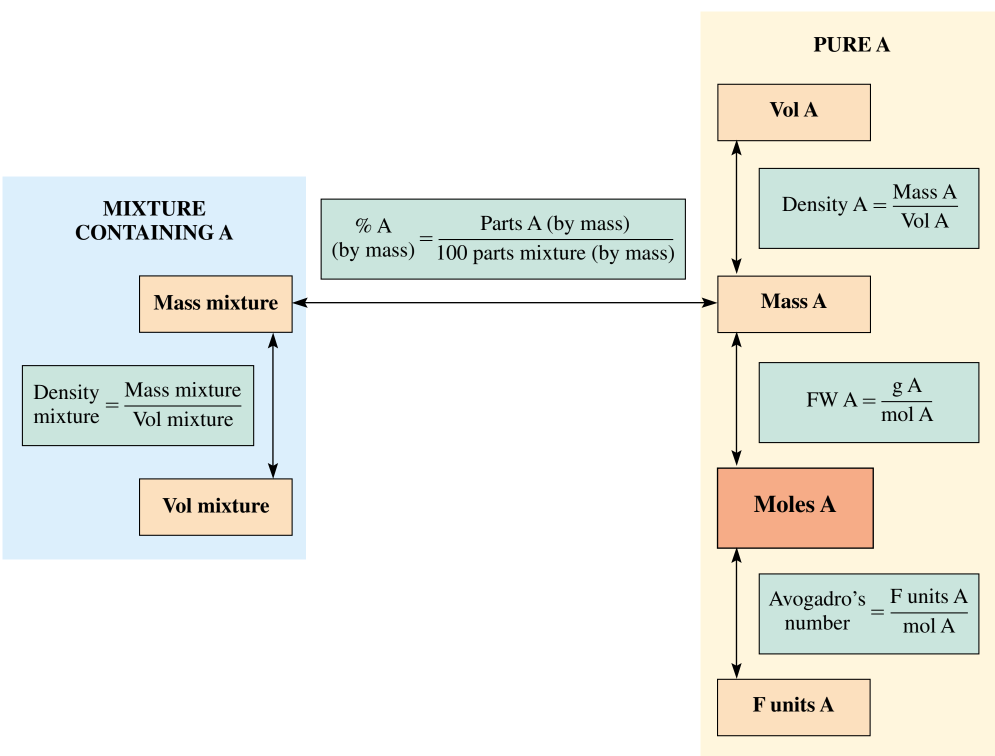


Figure 2-12 Some important relationships from Chapters 1 and 2. The relationships that provide unit factors are enclosed in green boxes.

Key Terms

Allotropic modifications (allotropes) Different forms of the same element in the same physical state.

Anhydrous Without water.

Anion An ion with a negative charge.

Atom The smallest particle of an element that maintains its chemical identity through all chemical and physical changes.

Atomic mass unit (amu) One twelfth of the mass of an atom of the carbon-12 isotope; a unit used for stating atomic and formula weights.

Atomic number The number of protons in the nucleus of an atom.

Atomic weight Weighted average of the masses of the constituent isotopes of an element; the relative mass of atoms of different elements.

Avogadro's number 6.022×10^{23} units of a specified item. See *Mole.*

Cation An ion with a positive charge.

Chemical formula Combination of symbols that indicates the chemical composition of a substance.

Composition stoichiometry Describes the quantitative (mass) relationships among elements in compounds.

Empirical formula See *Simplest formula.*

Formula Combination of symbols that indicates the chemical composition of a substance.

Formula unit The smallest repeating unit of a substance—for non-ionic substances, the molecule.

Formula weight The mass, in atomic mass units, of one formula unit of a substance. Numerically equal to the mass, in grams, of one mole of the substance (see *Molar mass*). This number is obtained by adding the atomic weights of the atoms specified in the formula.

Hydrate A crystalline sample that contains water, H_2O, and another compound in a fixed mole ratio. Examples include $CuSO_4 \cdot 5H_2O$ and $(COOH)_2 \cdot 2H_2O$.

Ion An atom or group of atoms that carries an electric charge. A positive ion is a *cation;* a negative ion is an *anion.*

Ionic compound A compound that is composed of cations and anions. An example is sodium chloride, NaCl.

Law of Constant Composition See *Law of Definite Proportions.*

Law of Definite Proportions Different samples of a pure compound always contain the same elements in the same proportions by mass; this corresponds to atoms of these elements in fixed numerical ratios. Also known as the Law of Constant Composition.

Law of Multiple Proportions When two elements, A and B, form more than one compound, the ratio of the masses of element B that combine with a given mass of element A in each of the compounds can be expressed by small whole numbers.

Molar mass The mass of substance in one mole of the substance; numerically equal to the formula weight of the substance. See *Formula weight;* see *Molecular weight.*

Mole 6.022×10^{23} (Avogadro's number of) formula units (or molecules, for a molecular substance) of the substance under discussion. The mass of one mole, in grams, is numerically equal to the formula (molecular) weight of the substance.

Molecular formula A formula that indicates the actual number of atoms present in a molecule of a molecular substance. Compare with *Simplest formula.*

Molecular weight The mass, in atomic mass units, of one molecule of a nonionic (molecular) substance. Numerically equal to the mass, in grams, of one mole of such a substance. This number is obtained by adding the atomic weights of the atoms specified in the formula.

Molecule The smallest particle of an element or compound that can have a stable independent existence.

Percent composition The mass percentage of each element in a compound.

Percent purity The mass percentage of a specified compound or element in an impure sample.

Polyatomic Consisting of more than one atom. Elements such as Cl_2, P_4, and S_8 exist as polyatomic molecules. Examples of polyatomic ions are the ammonium ion, NH_4^+, and the sulfate ion, SO_4^{2-}.

Simplest formula The smallest whole-number ratio of atoms present in a compound; also called empirical formula. Compare with *Molecular formula.*

Stoichiometry Description of the quantitative relationships among elements in compounds (composition stoichiometry) and among substances as they undergo chemical changes (reaction stoichiometry).

Structural formula A representation that shows how atoms are connected in a compound.

Exercises

Basic Ideas

1. (a) What is the origin of the word "stoichiometry"? (b) Distinguish between composition stoichiometry and reaction stoichiometry.

2. List the basic ideas of Dalton's atomic theory.

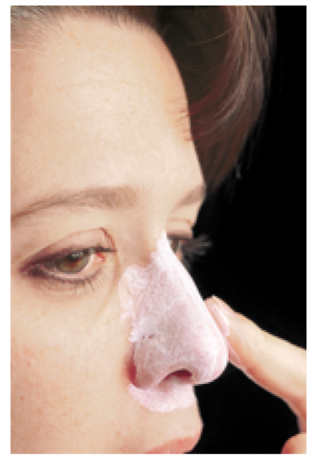
3. Give examples of molecules that contain (a) two atoms; (b) three atoms; (c) four atoms; (d) eight atoms.

4. Give the formulas of two diatomic molecules, a triatomic molecule, and two more complex molecules. Label each formula as being either the formula of an element or of a compound.

5. Which of the formulas you selected for Exercise 4 represent allotropes? If none, give two examples that are allotropes. Select a different element for each example.
6. Which of the compounds in Table 2-2 are inorganic compounds?
7. When can we correctly use the terms "formula weight" and "molecular weight" interchangeably?
8. What structural feature distinguishes organic compounds from inorganic compounds?

Names and Formulas

9. Write formulas for the following compounds: (a) nitric acid; (b) methyl alcohol; (c) sulfur dioxide; (d) acetic acid; (e) butane.
10. Name the following compounds: (a) H_2SO_4; (b) C_3H_8; (c) NH_3; (d) CH_3CH_2OH.
11. Name each of the following ions. Classify each as a monatomic or polyatomic ion. Classify each as a cation or an anion. (a) Na^+; (b) OH^-; (c) $SO_4{}^{2-}$; (d) S^{2-}; (e) Zn^{2+}; (f) Fe^{2+}.
12. Write the chemical symbol for each of the following ions. Classify each as a monatomic or polyatomic ion. Classify each as a cation or an anion. (a) potassium ion; (b) sulfate ion; (c) copper(II) ion; (d) ammonium ion; (e) carbonate ion.
13. Name each of the following compounds: (a) $MgCl_2$; (b) $Cu(NO_3)_2$; (c) Li_2SO_4; (d) $Ca(OH)_2$; (e) $FeSO_4$.
14. Write the chemical formula for each of the following ionic compounds: (a) potassium acetate; (b) ammonium sulfate; (c) zinc phosphate; (d) calcium oxide; (e) aluminum sulfide.
15. Write the chemical formula for the ionic compound formed between each of the following pairs of ions. Name each compound. (a) Na^+ and S^{2-}; (b) Al^{3+} and $SO_4{}^{2-}$; (c) Na^+ and $PO_4{}^{3-}$; (d) Mg^{2+} and $NO_3{}^-$; (e) Fe^{3+} and $CO_3{}^{2-}$.
16. Write the chemical formula for the ionic compound formed between each of the following pairs of ions. Name each compound. (a) Cu^{2+} and $CO_3{}^{2-}$; (b) Mg^{2+} and Cl^-; (c) $NH_4{}^+$ and $CO_3{}^{2-}$; (d) Zn^{2+} and OH^-; (e) Fe^{2+} and CH_3COO^-.
17. Define and illustrate the following: (a) ion; (b) cation; (c) anion; (d) polyatomic ion; (e) molecule.
18. (a) There are no *molecules* in ionic compounds. Why not? (b) What is the difference between a formula unit of an ionic compound and a polyatomic molecule?
19. Convert each of the following into a correct formula represented with correct notation. (a) $AlOH_3$; (b) Mg_2CO_3; (c) $Zn(CO_3)_2$; (d) $(NH_4)^2SO_4$; (e) $Mg_2(SO_4)_2$.
20. Write the formula of the compound produced by the combination of each of the following pairs of elements. Name each compound. (a) potassium and chlorine; (b) magnesium and chlorine; (c) sulfur and oxygen; (d) calcium and oxygen; (e) sodium and sulfur; (f) aluminum and sulfur.
21. Write the chemical formula of each of the following: (a) calcium carbonate—major component of coral, seashells, and limestone—found in antacid preparations; (b) magnesium sulfate—found in Epsom salts; (c) acetic acid—the acid in vinegar; (d) sodium hydroxide—common name is lye; (e) zinc oxide—used to protect from sunlight's UV rays when blended in an ointment.



Zinc oxide used as a sunscreen.

Atomic and Formula Weights

22. (a) What is the atomic weight of an element? (b) Why can atomic weights be referred to as relative numbers?
23. (a) What is the atomic mass unit (amu)? (b) The atomic weight of vanadium is 50.942 amu, and the atomic weight of ruthenium is 101.07 amu. What can we say about the relative masses of V and Ru atoms?
24. What is the mass ratio (four significant figures) of one atom of Rb to one atom of Cl?
25. A sample of 6.68 g of calcium combines exactly with 6.33 g of fluorine, forming calcium fluoride, CaF_2. Find the relative masses of the atoms of calcium and fluorine. Check your answer using a table of atomic weights. If the formula were not known, could you still do this calculation?
26. Determine the formula weight of each of the following substances: (a) bromine, Br_2; (b) water, H_2O; (c) saccharin, $C_7H_5NSO_3$; (d) potassium dichromate, $K_2Cr_2O_7$.
27. Determine the formula weight of each of the following substances: (a) calcium sulfate, $CaSO_4$; (b) butane, C_4H_{10}; (c) the sulfa drug sulfanilamide, $C_6H_4SO_2(NH_2)_2$; (d) uranyl phosphate, $(UO_2)_3(PO_4)_2$.
28. Determine the formula weight of each of the following common acids: (a) hydrochloric acid, HCl; (b) nitric acid, HNO_3; (c) phosphoric acid, H_3PO_4; (d) sulfuric acid, H_2SO_4.

The Mole Concept

29. A large neon sign is to be filled with a mixture of gases, including 8.575 g neon. What number of moles is this?

30. How many molecules are in 18.0 g of each of the following substances? (a) CO; (b) N_2; (c) P_4; (d) P_2. (e) Do parts (c) and (d) contain the same number of atoms of phosphorus?

31. Sulfur molecules exist under various conditions as S_8, S_6, S_4, S_2, and S. (a) Is the mass of one mole of each of these molecules the same? (b) Is the number of molecules in one mole of each of these molecules the same? (c) Is the mass of sulfur in one mole of each of these molecules the same? (d) Is the number of atoms of sulfur in one mole of each of these molecules the same?

32. How many moles of substance are contained in each of the following samples? (a) 18.3 g of NH_3; (b) 5.32 g of ammonium bromide; (c) 6.6 g of PCl_5; (d) 215 g of Sn.

33. How many moles of substance are contained in each of the following samples? (a) 36.2 g of diethyl ether; (b) 15.6 g of calcium carbonate; (c) 16.7 g of acetic acid; (d) 19.3 g of ethanol.

34. Complete the following table. Refer to a table of atomic weights.

Element	Atomic Weight	Mass of One Mole of Atoms
(a) Mg	________	________
(b) ______	79.904 amu	________
(c) Cl	________	________
(d) ______	________	51.9961 g

35. Complete the following table. Refer to a table of atomic weights.

Element	Formula	Mass of One Mole of Molecules
(a) Br	Br_2	________
(b) ______	H_2	________
(c) ______	P_4	________
(d) ______	______	20.1797 g
(e) S	______	256.528 g
(f) O	______	________

36. Complete the following table.

Moles of Compound	Moles of Cations	Moles of Anions
1 mol KCl	________	________
2 mol Na_2SO_4	________	________
0.2 mol calcium nitrate	________	________
________	0.50 mol NH_4^+	0.25 mol SO_4^{2+}

37. What mass, in grams, should be weighed for an experiment that requires 1.54 mol of $(NH_4)_2HPO_4$?

38. How many hydrogen atoms are contained in 125 grams of propane, C_3H_8?

39. How many atoms of C, H, and O are in each of the following? (a) 1.24 mol of glucose, $C_6H_{12}O_6$; (b) 3.31×10^{19} glucose ($C_6H_{12}O_6$) molecules; (c) 0.275 g of glucose.

40. Calculate the mass in grams and kilograms of 1.458 moles of gold.

41. An atom of an element has a mass ever-so-slightly greater than twice the mass of an Ni atom. Identify the element.

42. Calculate the number of Ni atoms in 1.0 millionth of a gram of nickel.

43. Calculate the number of Ni atoms in 1.0 trillionth of a gram of nickel.

44. What is the mass of 10.0 million methane, CH_4, molecules?

45. A sample of ethane, C_2H_6, has the same mass as 10.0 million molecules of methane, CH_4. How many C_2H_6 molecules does the sample contain?

Percent Composition

46. Calculate the percent composition of each of the following compounds: (a) nicotine, $C_{10}H_{14}N_2$; (b) vitamin E, $C_{29}H_{50}O_2$; (c) vanillin, $C_8H_8O_3$.

47. Calculate the percent composition of each of the following compounds: (a) menthol, $C_{10}H_{19}OH$; (b) carborundum, SiC; (c) aspirin, $C_9H_8O_4$.

48. Calculate the percent by mass of silver found in a particular mineral that is determined to be silver carbonate.

49. What percent by mass of iron(II) phosphate is iron?

***50.** Copper is obtained from ores containing the following minerals: azurite, $Cu_3(CO_3)_2(OH)_2$; chalcocite, Cu_2S; chalcopyrite, $CuFeS_2$; covelite, CuS; cuprite, Cu_2O; and malachite, $Cu_2CO_3(OH)_2$. Which mineral has the highest copper content as a percent by mass?

Determination of Simplest and Molecular Formulas

51. Determine the simplest formula for each of the following compounds: (a) copper(II) tartrate: 30.03% Cu; 22.70% C; 1.91% H; 45.37% O. (b) nitrosyl fluoroborate: 11.99% N; 13.70% O; 9.25% B; 65.06% F.

52. The hormone norepinephrine is released in the human body during stress and increases the body's metabolic rate. Like many biochemical compounds, norepinephrine is composed of carbon, hydrogen, oxygen, and nitrogen. The percent composition of this hormone is 56.8% C, 6.56% H, 28.4% O, and 8.28% N. What is the simplest formula of norepinephrine?

53. (a) A sample of a compound is found to contain 5.60 g N, 14.2 g Cl, and 0.800 g H. What is the simplest formula of this compound? (b) A sample of another compound containing the same elements is found to be 26.2% N, 66.4%

Cl, and 7.5% H. What is the simplest formula of this compound?

54. A common product found in nearly every kitchen contains 27.37% sodium, 1.20% hydrogen, 14.30% carbon, and 57.14% oxygen. The simplest formula is the same as the formula of the compound. Find the formula of this compound.

55. Bupropion is present in a medication that is an antidepressant and is also used to aid in quitting smoking. The composition of bupropion is 65.13% carbon, 7.57% hydrogen, 14.79% chlorine, 5.84% nitrogen, and 6.67% oxygen. The simplest formula is the same as the molecular formula of this compound. Determine the formula.

56. Lysine is an essential amino acid. One experiment showed that each molecule of lysine contains two nitrogen atoms. Another experiment showed that lysine contains 19.2% N, 9.64% H, 49.3% C, and 21.9% O by mass. What is the molecular formula for lysine?

57. A 2.00-g sample of a compound gave 4.86 g of CO_2 and 2.03 g of H_2O on combustion in oxygen. The compound is known to contain only C, H, and O. What is its simplest formula?

58. A 0.1647-gram sample of a pure hydrocarbon was burned in a C-H combustion train to produce 0.5694 gram of CO_2 and 0.0826 gram of H_2O. Determine the masses of C and H in the sample and the percentages of these elements in this hydrocarbon.

59. Naphthalene is a hydrocarbon that is used for mothballs. A 0.3204-gram sample of naphthalene was burned in a C-H combustion train to produce 1.100 grams of carbon dioxide and 0.1802 grams of water. What masses and percentages of C and H are present in naphthalene?

60. Combustion of 0.5707 mg of a hydrocarbon produces 1.790 mg of CO_2. What is the simplest formula of the hydrocarbon?

***61.** Complicated chemical reactions occur at hot springs on the ocean floor. One compound obtained from such a hot spring consists of Mg, Si, H, and O. From a 0.334-g sample, the Mg is recovered as 0.115 g of MgO; H is recovered as 25.7 mg of H_2O; and Si is recovered as 0.172 g of SiO_2. What is the simplest formula of this compound?

62. A 1.000-gram sample of an alcohol was burned in oxygen to produce 1.913 g of CO_2 and 1.174 g of H_2O. The alcohol contained only C, H, and O. What is the simplest formula of the alcohol?

63. An alcohol is 64.81% C, 13.60% H, and 21.59% O by mass. Another experiment shows that its molecular weight is approximately 74 amu. What is the molecular formula of the alcohol?

64. Skatole is found in coal tar and in human feces. It contains three elements: C, H, and N. It is 82.40% C and 6.92% H by mass. Its simplest formula is its molecular formula. What are (a) the formula and (b) the molecular weight of skatole?

65. Testosterone, the male sex hormone, contains only C, H, and O. It is 79.12% C and 9.79% H by mass. Each molecule contains two O atoms. What are (a) the molecular weight and (b) the molecular formula for testosterone?

***66.** The beta-blocker drug, timolol, is expected to reduce the need for heart bypass surgery. Its composition by mass is 49.4% C, 7.64% H, 17.7% N, 15.2% O, and 10.1% S. The mass of 0.0100 mol of timolol is 3.16 g. (a) What is the simplest formula of timolol? (b) What is the molecular formula of timolol?

The Law of Multiple Proportions

67. Show that the compounds water, H_2O, and hydrogen peroxide, H_2O_2, obey the Law of Multiple Proportions.

68. Nitric oxide, NO, is produced in internal combustion engines. When NO comes in contact with air, it is quickly converted into nitrogen dioxide, NO_2, a very poisonous, corrosive gas. What mass of O is combined with 3.00 g of N in (a) NO and (b) NO_2? Show that NO and NO_2 obey the Law of Multiple Proportions.

69. Sulfur forms two chlorides. A 30.00-gram sample of one chloride decomposes to give 5.53 g of S and 24.47 g of Cl. A 30.00-gram sample of the other chloride decomposes to give 3.93 g of S and 26.07 g of Cl. Show that these compounds obey the Law of Multiple Proportions.

70. What mass of oxygen is combined with 3.65 g of sulfur in (a) sulfur dioxide, SO_2, and in (b) sulfur trioxide, SO_3?

Interpretation of Chemical Formulas

71. One prominent ore of copper contains chalcopyrite, $CuFeS_2$. How many pounds of copper are contained in 2.63 pounds of pure $CuFeS_2$?

72. Mercury occurs as a sulfide ore called *cinnabar*, HgS. How many grams of mercury are contained in 887 g of pure HgS?

73. (a) How many grams of copper are contained in 325 g of $CuSO_4$? (b) How many grams of copper are contained in 325 g of $CuSO_4 \cdot 5H_2O$?

74. What mass of $KMnO_4$ would contain 15.0 g of manganese?

75. What mass of azurite, $Cu_3(CO_3)_2(OH)_2$, would contain 610 g of copper?

76. Two minerals that contain copper are chalcopyrite, $CuFeS_2$, and chalcocite, Cu_2S. What mass of chalcocite would contain the same mass of copper as is contained in 125 pounds of chalcopyrite?

77. Tungsten is a very dense metal (19.3 g/cm^3) with extremely high melting and boiling points (3370°C and 5900°C). When a small amount of it is included in steel, the resulting alloy is far harder and stronger than ordinary steel. Two important ores of tungsten are $FeWO_4$ and $CaWO_4$. How many grams of $CaWO_4$ would contain the same mass of tungsten that is present in 569 g of $FeWO_4$?

***78.** When a mole of $CuSO_4 \cdot 5H_2O$ is heated to 110°C, it loses four moles of H_2O to form $CuSO_4 \cdot H_2O$. When it is heated to temperatures above 150°C, the other mole of H_2O is lost. (a) How many grams of $CuSO_4 \cdot H_2O$ could be obtained by heating 695 g of $CuSO_4 \cdot 5H_2O$ to 110°C? (b) How many grams of anhydrous $CuSO_4$ could be obtained by heating 695 g of $CuSO_4 \cdot 5H_2O$ to 180°C?

Percent Purity

79. A particular ore of lead, galena, is 10.0% lead sulfide, PbS, and 90.0% impurities by weight. What mass of lead is contained in 50.0 grams of this ore?

80. What mass of chromium is present in 150 grams of an ore of chromium that is 65.0% chromite, $FeCr_2O_4$, and 35.0% impurities by mass? If 90.0% of the chromium can be recovered from 100.0 grams of the ore, what mass of pure chromium is obtained?

81. What masses of (a) Sr and (b) N are contained in 106.7 g of 88.2% pure $Sr(NO_3)_2$? Assume that the impurities do not contain the elements mentioned.

82. (a) What weight of magnesium carbonate is contained in 315 pounds of an ore that is 27.7% magnesium carbonate by weight? (b) What weight of impurities is contained in the sample? (c) What weight of magnesium is contained in the sample? (Assume that no magnesium is present in the impurities.)

83. Vinegar is 5.0% acetic acid, $C_2H_4O_2$, by mass. (a) How many grams of acetic acid are contained in 24.0 g of vinegar? (b) How many pounds of acetic acid are contained in 24.0 pounds of vinegar? (c) How many grams of sodium chloride, NaCl, are contained in 24.0 g of saline solution that is 5.0% NaCl by mass?

***84.** What is the percent by mass of copper sulfate, $CuSO_4$, in a sample of copper sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$? (b) What is the percent by mass of $CuSO_4$ in a sample that is 72.4% $CuSO_4 \cdot 5H_2O$ by mass?

Mixed Examples

85. How many moles of chlorine atoms are contained in each of the following? (a) 35.45×10^{23} Cl atoms; (b) 35.45×10^{23} Cl_2 molecules; (c) 35.45 g of chlorine; (d) 35.45 mol of Cl_2.

86. What is the *maximum* number of moles of CO_2 that could be obtained from the carbon in each of the following? (a) 4.00 mol of $Ru_2(CO_3)_3$; (b) 4.00 mol of $CaCO_3$; (c) 4.00 mol of $Co(CO)_6$.

87. (a) How many formula units are contained in 154.3 g of K_2MoO_4? (b) How many potassium ions? (c) How many MoO_4^{2-} ions? (d) How many atoms of all kinds?

88. (a) How many moles of ozone molecules are contained in 64.0 g of ozone, O_3? (b) How many moles of oxygen atoms are contained in 64.0 g of ozone? (c) What mass of O_2 would contain the same number of oxygen atoms as 64.0 g of ozone? (d) What mass of oxygen gas, O_2, would contain the same number of molecules as 64.0 g of ozone?

89. Cocaine has the following percent composition by mass: 67.30% C, 6.930% H, 21.15% O, and 4.62% N. What is the simplest formula of cocaine?

90. A compound with the molecular weight of 56.0 g was found as a component of photochemical smog. The compound is composed of carbon and oxygen, 42.9% and 57.1%, respectively. What is the formula of this compound?

91. A carbon-hydrogen-oxygen compound, MW = 90 g, is analyzed and found to be 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen. What is the formula of this compound?

92. Find the number of moles of Ag needed to form each of the following: (a) 0.235 mol Ag_2S; (b) 0.235 mol Ag_2O; (c) 0.235 g Ag_2S; (d) 2.35×10^{20} formula units of Ag_2S.

93. A metal, M, forms an oxide having the simplest formula M_2O_3. This oxide contains 52.9% of the metal by mass. (a) Calculate the atomic weight of the metal. (b) Identify the metal.

94. Three samples of magnesium oxide were analyzed to determine the mass ratios O/Mg, giving the following results:

$$\frac{1.60 \text{ g O}}{2.43 \text{ g Mg}}, \quad \frac{0.658 \text{ g O}}{1.00 \text{ g Mg}}, \quad \frac{2.29 \text{ g O}}{3.48 \text{ g Mg}}$$

Which law of chemical combination is illustrated by these data?

***95.** The molecular weight of hemoglobin is about 65,000 g/mol. Hemoglobin contains 0.35% Fe by mass. How many iron atoms are in a hemoglobin molecule?

***96.** More than 1 billion pounds of adipic acid (MW 146.1 g/mol) is manufactured in the United States each year. Most of it is used to make synthetic fabrics. Adipic acid contains only C, H, and O. Combustion of a 1.6380-g sample of adipic acid gives 2.960 g of CO_2 and 1.010 g of H_2O. (a) What is the simplest formula for adipic acid? (b) What is its molecular formula?

97. A filled 25-L container of an unknown, unlabeled liquid was found in a storeroom and had to be identified to determine a method of disposal. The compound was found to contain only hydrogen and carbon. A 1.750-g sample of the compound was burned in a pure oxygen atmosphere; 1.211 g H_2O and 5.916 g CO_2 were collected. Determine the simplest formula.

98. An unknown sample weighing 1.50 g was found to contain only manganese and sulfur. The sample was completely reacted with oxygen and produced 1.22 g manganese(II) oxide, MnO, and 1.38 g sulfur trioxide. What is the simplest formula for this compound?

99. Copper(II) sulfate exists as a baby-blue powder when anhydrous and as a deep blue crystal when hydrated with five water molecules, $CuSO_4 \cdot 5H_2O$. Which of these two compounds contains more copper per mole of compound?

What is the ratio of percent by mass of copper in the anhydrous compound to the percent by mass of copper in the hydrated compound?

CONCEPTUAL EXERCISES

100. What mass of NaCl would contain the same *total* number of ions as 245 g of $MgCl_2$?

101. How many atoms of oxygen are in 17.9325 g of sulfuric acid?

102. In the "button" and "button hole" analogy of writing formulas for ionic compounds, one may think of positive charges as buttons and negative charges as button holes. One prepares formulas by combining the buttons (positive charges) with an equal number of button holes (negative charges) so that every button will be associated with a single buttonhole and vice versa. Using this analogy, how many buttons (positive charges) are associated with a single cation (positive ion) in each of the following ionic compounds? (a) NaCl; (b) Na_2SO_4; (c) $CaSO_4$; (d) $Al_2(SO_4)_3$.

103. Two deposits of minerals containing silver are found. One of the deposits contains silver oxide, and the other contains silver sulfide. The deposits can be mined at the same price per ton of the original silver-containing compound, but only one deposit can be mined by your company. Which of the deposits would you recommend and why?

104. A decision is to be made as to the least expensive source of zinc. One source of zinc is zinc sulfate, $ZnSO_4$, and another is zinc acetate dihydrate, $Zn(CH_3COO)_2 \cdot 2H_2O$. These two sources of zinc can be purchased at the same price per kilogram of compound. Which is the most economical source of zinc and by how much?

105. Assume that a penny is $\frac{1}{16}$ in. thick and that the moon is 222,000 mi at its closest approach to the earth (perigee). Show by calculation whether or not a picomole of pennies stacked on their faces would reach from the earth to the moon.

BUILDING YOUR KNOWLEDGE

NOTE *Beginning with this chapter, exercises under the "Building Your Knowledge" heading will often require that you use skills, concepts, or information that you should have mastered in earlier chapters. This provides you an excellent opportunity to "tie things together" as you study.*

106. A 22-mL (19-g) sample of an unknown liquid is analyzed, and the percent composition is found to be 53% C, 11% H, and 36% O. Is this compound likely to be ethanol, CH_3CH_2OH? Give two reasons for your answer. (*Hint:* See Table 1-8.)

107. Three allotropes of phosphorus are known, with molecular weights of 62.0, 31.0, and 124.0 amu, respectively. Write the molecular formula for each allotrope.

108. Near room temperature, the density of water is 1.00 g/mL, and the density of ethanol (grain alcohol) is 0.789 g/mL. What volume of ethanol contains the same number of molecules as are present in 175 mL of H_2O?

109. Calculate the volume of 2.00 mol of mercury, a liquid metal. (*Hint:* See Table 1-8.)

110. In Chapter 1 you learned that the specific heat of water is $4.18\ J/g \cdot °C$. The molar heat capacity is defined as the specific heat or heat capacity per mole of material. Calculate the molar heat capacity for water. What value(s) limited the number of significant figures in your answer?