

# *Chemical Formulas and Chemical Compounds*

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*Chemists use chemical names and formulas to describe the atomic composition of compounds.*

# Chemical Names and Formulas

## SECTION 7-1

### OBJECTIVES

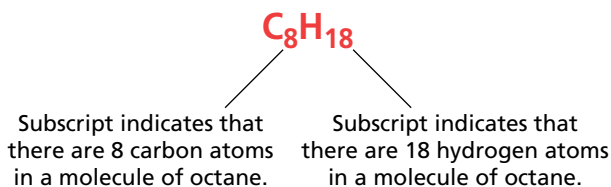
- Explain the significance of a chemical formula.
- Determine the formula of an ionic compound formed between two given ions.
- Name an ionic compound given its formula.
- Using prefixes, name a binary molecular compound from its formula.
- Write the formula of a binary molecular compound given its name.

The total number of natural and synthetic chemical compounds runs in the millions. For some of these substances, certain common names remain in everyday use. For example, calcium carbonate is better known as limestone, and sodium chloride is usually referred to simply as table salt. And everyone recognizes dihydrogen monoxide by its popular name, water.


Unfortunately, common names usually give no information about chemical composition. To describe the atomic makeup of compounds, chemists use systematic methods for naming compounds and for writing chemical formulas. In this chapter, you will be introduced to some of the rules used to identify simple chemical compounds.


### Significance of a Chemical Formula

Recall that a chemical formula indicates the relative number of atoms of each kind in a chemical compound. For a molecular compound, the chemical formula reveals the number of atoms of each element contained in a single molecule of the compound, as shown below for the hydrocarbon octane. (*Hydrocarbons* are molecular compounds composed solely of carbon and hydrogen.)

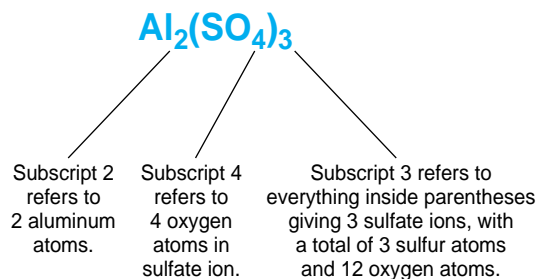


Unlike a molecular compound, an ionic compound consists of a lattice of positive and negative ions held together by mutual attraction. The chemical formula for an ionic compound represents one formula unit—the simplest ratio of the compound's positive ions (cations) and its negative ions (anions). The chemical formula for aluminum sulfate, an ionic compound consisting of aluminum cations and polyatomic sulfate anions, is written as shown on page 204.

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Note how the parentheses are used. They surround the polyatomic anion to identify it as a unit. The subscript 3 refers to the entire unit. Notice also that there is no subscript written next to the symbol for sulfur. When there is no subscript written next to an atom's symbol, the value of the subscript is understood to be 1.

## Monatomic Ions

By gaining or losing electrons, many main-group elements form ions with noble-gas configurations. For example, Group 1 metals lose one electron to give 1+ cations, such as  $\text{Na}^+$ . Group 2 metals lose two electrons to give 2+ cations, such as  $\text{Mg}^{2+}$ . *Ions formed from a single atom are known as **monatomic ions**.* The nonmetals of Groups 15, 16, and 17 gain electrons to form anions. For example, in ionic compounds nitrogen forms the 3− anion,  $\text{N}^{3-}$ . The three added electrons plus the five outermost electrons in nitrogen atoms give a completed outermost octet. Similarly, the Group 16 elements oxygen and sulfur form 2− anions, and the Group 17 halogens form 1− anions.

Not all main-group elements readily form ions, however. Rather than gain or lose electrons, atoms of carbon and silicon form covalent bonds in which they share electrons with other atoms. Other elements tend to form ions that do not have noble-gas configurations. For instance, it is difficult for the Group 14 metals tin and lead to lose four electrons to achieve a noble-gas configuration. Instead, they tend to lose the two electrons in their outer *p* orbitals but retain the two electrons in their outer *s* orbitals to form 2+ cations. (Tin and lead can also form molecular compounds in which all four valence electrons are involved in covalent bonding.)

Elements from the *d*-block form 2+, 3+, or, in a few cases, 1+ or 4+ cations. Many *d*-block elements form two ions of different charges. For example, copper forms 1+ cations and 2+ cations. Iron and chromium, on the other hand, each form 2+ cations as well as 3+ cations. And vanadium and lead form 2+, 3+, and 4+ cations.

## Naming Monatomic Ions

Monatomic cations are identified simply by the element's name, as illustrated by the examples at left. Naming monatomic anions is slightly more

### Naming cations



Potassium cation



Magnesium cation

complicated. First, the ending of the element's name is dropped. Then the ending *-ide* is added to the root name, as illustrated by the examples at right.

The names and symbols of the common monatomic cations and anions are organized according to their charges in Table 7-1. The names of many of the ions in the table include Roman numerals. These numerals are part of the *Stock system* of naming chemical ions and elements. You will read more about the Stock system and other systems of naming chemicals later in this chapter.

### Naming anions

<b>F</b>	<b>F<sup>-</sup></b>
Fluorine	Fluoride anion
<b>N</b>	<b>N<sup>3-</sup></b>
Nitrogen	Nitride anion

**TABLE 7-1 Some Common Monatomic Ions**

Main-group elements							
1+		2+		3+			
lithium	Li <sup>+</sup>						
sodium	Na <sup>+</sup>	magnesium	Mg <sup>2+</sup>	aluminum	Al <sup>3+</sup>		
potassium	K <sup>+</sup>	calcium	Ca <sup>2+</sup>				
rubidium	Rb <sup>+</sup>	strontium	Sr <sup>2+</sup>				
cesium	Cs <sup>+</sup>	barium	Ba <sup>2+</sup>				
1-		2-		3-			
fluoride	F <sup>-</sup>	oxide	O <sup>2-</sup>	nitride	N <sup>3-</sup>		
chloride	Cl <sup>-</sup>	sulfide	S <sup>2-</sup>				
bromide	Br <sup>-</sup>						
iodide	I <sup>-</sup>						
d-Block elements and others with multiple ions							
1+		2+		3+		4+	
copper(I)	Cu <sup>+</sup>	cadmium	Cd <sup>2+</sup>	chromium(III)	Cr <sup>3+</sup>	lead(IV)	Pb <sup>4+</sup>
silver	Ag <sup>+</sup>	chromium(II)	Cr <sup>2+</sup>	iron(III)	Fe <sup>3+</sup>	vanadium(IV)	V <sup>4+</sup>
		cobalt(II)	Co <sup>2+</sup>	vanadium(III)	V <sup>3+</sup>	tin(IV)	Sn <sup>4+</sup>
		copper(II)	Cu <sup>2+</sup>				
		iron(II)	Fe <sup>2+</sup>				
		lead(II)	Pb <sup>2+</sup>				
		manganese(II)	Mn <sup>2+</sup>				
		mercury(II)	Hg <sup>2+</sup>				
		nickel(II)	Ni <sup>2+</sup>				
		tin(II)	Sn <sup>2+</sup>				
		vanadium(II)	V <sup>2+</sup>				
		zinc	Zn <sup>2+</sup>				

## Binary Ionic Compounds

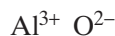
Compounds composed of two different elements are known as **binary compounds**. In a binary ionic compound, the total numbers of positive charges and negative charges must be equal. Therefore, the formula for such a compound can be written given the identities of the compound's ions. For example, magnesium and bromine combine to form the ionic compound magnesium bromide. Magnesium, a Group 2 metal, forms the  $\text{Mg}^{2+}$  cation. Note that the  $^{2+}$  in  $\text{Mg}^{2+}$  is written as a superscript. Bromine, a halogen, forms the  $\text{Br}^-$  anion when combined with a metal. In each formula unit of magnesium bromide, two  $\text{Br}^-$  anions are required to balance the  $2+$  charge of the  $\text{Mg}^{2+}$  cation. The compound's formula must therefore indicate one  $\text{Mg}^{2+}$  cation and two  $\text{Br}^-$  anions. The symbol for the cation is written first.

*Ions combined:*  $\text{Mg}^{2+}$ ,  $\text{Br}^-$ ,  $\text{Br}^-$     *Chemical formula:*  $\text{MgBr}_2$

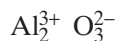
Note that the  $_2$  in  $\text{Br}_2$  is written as a subscript. The charges of the ions are not included in the formula. This is usually the case when writing formulas for binary ionic compounds.

As an aid to determining subscripts in formulas for ionic compounds, the positive and negative charges can be “crossed over.” Crossing over is a method of balancing the charges between ions in an ionic compound. For example, the formula for the compound formed by the aluminum ion,  $\text{Al}^{3+}$ , and the oxide ion,  $\text{O}^{2-}$ , is determined as follows.

1. Write the symbols for the ions side by side. Write the cation first.



2. Cross over the charges by using the absolute value of each ion's charge as the subscript for the other ion.



3. Check the subscripts and divide them by their largest common factor to give the smallest possible whole-number ratio of ions. Then write the formula.

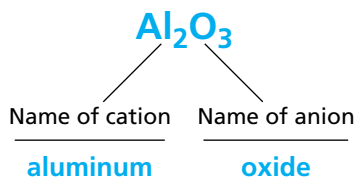
Multiplying the charge by the subscript shows that the charge on two  $\text{Al}^{3+}$  cations ( $2 \times 3+ = 6+$ ) equals the charge on three  $\text{O}^{2-}$  anions ( $3 \times 2- = 6-$ ). The largest common factor of the subscripts is 1. The correct formula is therefore written as follows.



## Naming Binary Ionic Compounds

The **nomenclature**, or *naming system*, of binary ionic compounds involves combining the names of the compound's positive and negative

ions. The name of the cation is given first, followed by the name of the anion. For most simple ionic compounds, the ratio of the ions is not indicated in the compound's name because it is understood based on the relative charges of the compound's ions. The naming of a simple binary ionic compound is illustrated below.



### SAMPLE PROBLEM 7-1

**Write the formulas for the binary ionic compounds formed between the following elements:**

a. zinc and iodine

b. zinc and sulfur

#### SOLUTION

*Write the symbols for the ions side by side. Write the cation first.*

a.  $\text{Zn}^{2+} \text{I}^-$

b.  $\text{Zn}^{2+} \text{S}^{2-}$

*Cross over the charges to give subscripts.*

a.  $\text{Zn}_1^{2+} \text{I}_2^-$

b.  $\text{Zn}_2^{2+} \text{S}_2^{2-}$

*Check the subscripts and divide them by their largest common factor to give the smallest possible whole-number ratio of ions. Then write the formula.*

a. The subscripts are mathematically correct because they give equal total charges of  $1 \times 2+ = 2+$  and  $2 \times 1- = 2-$ . The largest common factor of the subscripts is 1.

The smallest possible whole-number ratio of ions in the compound is therefore 1:2.

The subscript 1 is not written, so the formula is  $\text{ZnI}_2$ .

b. The subscripts are mathematically correct because they give equal total charges of  $2 \times 2+ = 4+$  and  $2 \times 2- = 4-$ . The largest common factor of the subscripts is 2.

The smallest whole-number ratio of ions in the compound is therefore 1:1. The correct formula is  $\text{ZnS}$ .

#### PRACTICE

1. Write formulas for the binary ionic compounds formed between the following elements:

a. potassium and iodine

d. aluminum and sulfur

b. magnesium and chlorine

e. aluminum and nitrogen

c. sodium and sulfur

*Answer*

1. a. KI d.  $\text{Al}_2\text{S}_3$

b.  $\text{MgCl}_2$  e.  $\text{AlN}$

c.  $\text{Na}_2\text{S}$

2. Name the binary ionic compounds indicated by the following formulas:

a.  $\text{AgCl}$

e.  $\text{BaO}$

b.  $\text{ZnO}$

f.  $\text{CaCl}_2$

c.  $\text{CaBr}_2$

d.  $\text{SrF}_2$

2. a. silver chloride

b. zinc oxide

c. calcium bromide

d. strontium fluoride

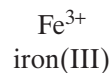
e. barium oxide

f. calcium chloride

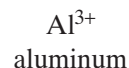


## The Stock System of Nomenclature

Some elements, such as iron, form two or more cations with different charges. To distinguish the ions formed by such elements, the Stock system of nomenclature is used. This system uses a Roman numeral to indicate an ion's charge. The numeral is enclosed in parentheses and placed *immediately* after the metal name.

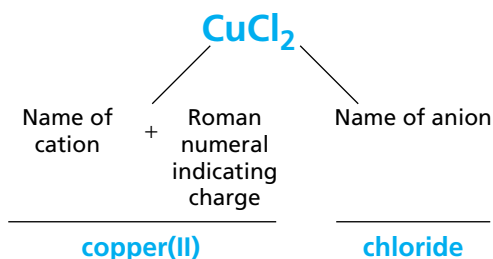


Names of metals that commonly form only one cation *do not* include a Roman numeral.



There is no element that commonly forms more than one monatomic anion.

Naming a binary ionic compound according to the Stock system is illustrated below.



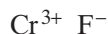
**FIGURE 7-1** Different cations of the same metal form different compounds even when they combine with the same anion. Compare (a) lead(IV) oxide,  $\text{PbO}_2$ , with (b) lead(II) oxide,  $\text{PbO}$ .

**SAMPLE PROBLEM 7-2**

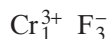
Write the formula and give the name for the compound formed by the ions  $\text{Cr}^{3+}$  and  $\text{F}^-$ .

**SOLUTION**

Write the symbols for the ions side by side. Write the cation first.



Cross over the charges to give subscripts.



Check the subscripts and write the formula.

The subscripts are correct because they give charges of  $1 \times 3+ = 3+$  and  $3 \times 1- = 3-$ . The largest common factor of the subscripts is 1, so the smallest whole-number ratio of the ions is 1:3. The formula is therefore  $\text{CrF}_3$ . As Table 7-1 on page 205 shows, chromium forms more than one ion. Therefore, the name of the  $3+$  chromium ion must be followed by a Roman numeral indicating its charge. The compound's name is chromium(III) fluoride.

**PRACTICE**

1. Write the formula and give the name for the compounds formed between the following ions:

- $\text{Cu}^{2+}$  and  $\text{Br}^-$
- $\text{Fe}^{2+}$  and  $\text{O}^{2-}$
- $\text{Pb}^{2+}$  and  $\text{Cl}^-$
- $\text{Hg}^{2+}$  and  $\text{S}^{2-}$
- $\text{Sn}^{2+}$  and  $\text{F}^-$
- $\text{Fe}^{3+}$  and  $\text{O}^{2-}$

2. Give the names for the following compounds:

- $\text{CuO}$
- $\text{CoF}_3$
- $\text{SnI}_4$
- $\text{FeS}$

Answer

- $\text{CuBr}_2$ , copper(II) bromide
  - $\text{FeO}$ , iron(II) oxide
  - $\text{PbCl}_2$ , lead(II) chloride
  - $\text{HgS}$ , mercury(II) sulfide
  - $\text{SnF}_2$ , tin(II) fluoride
  - $\text{Fe}_2\text{O}_3$ , iron(III) oxide
- copper(II) oxide
  - cobalt(III) fluoride
  - tin(IV) iodide
  - iron(II) sulfide

**Compounds Containing Polyatomic Ions**

Table 7-2 on page 210 lists some common polyatomic ions. All but the ammonium ion are negatively charged and most are **oxyanions**—polyatomic ions that contain oxygen. In several cases, two different oxyanions are formed by the same two elements. Nitrogen and oxygen, for example, are combined in both  $\text{NO}_2^-$  and  $\text{NO}_3^-$ . When naming compounds containing such oxyanions, the most common ion is given the ending *-ate*. The ion with one less oxygen atom is given the ending *-ite*.



Sometimes two elements form more than two different oxyanions. In this case, an anion with one less oxygen than the *-ite* anion is given the



**TABLE 7-2 Some Polyatomic Ions**

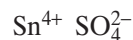
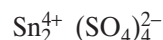
1+		2+			
ammonium	$\text{NH}_4^+$	dimercury*	$\text{Hg}_2^{2+}$		
1-		2-		3-	
acetate	$\text{CH}_3\text{COO}^-$	carbonate	$\text{CO}_3^{2-}$	phosphate	$\text{PO}_4^{3-}$
bromate	$\text{BrO}_3^-$	chromate	$\text{CrO}_4^{2-}$	arsenate	$\text{AsO}_4^{3-}$
chlorate	$\text{ClO}_3^-$	dichromate	$\text{Cr}_2\text{O}_7^{2-}$		
chlorite	$\text{ClO}_2^-$	hydrogen phosphate	$\text{HPO}_4^{2-}$		
cyanide	$\text{CN}^-$	oxalate	$\text{C}_2\text{O}_4^{2-}$		
dihydrogen phosphate	$\text{H}_2\text{PO}_4^-$	peroxide	$\text{O}_2^{2-}$		
hydrogen carbonate (bicarbonate)	$\text{HCO}_3^-$	sulfate	$\text{SO}_4^{2-}$		
hydrogen sulfate	$\text{HSO}_4^-$	sulfite	$\text{SO}_3^{2-}$		
hydroxide	$\text{OH}^-$				
hypochlorite	$\text{ClO}^-$				
nitrate	$\text{NO}_3^-$				
nitrite	$\text{NO}_2^-$				
perchlorate	$\text{ClO}_4^-$				
permanganate	$\text{MnO}_4^-$				

\*The mercury(I) cation exists as two  $\text{Hg}^+$  ions joined together by a covalent bond and is written as  $\text{Hg}_2^{2+}$ .

prefix *hypo*-. An anion with one more oxygen than the *-ate* anion is given the prefix *per*-. This nomenclature is illustrated by the four oxyanions formed between chlorine and oxygen.



Compounds containing polyatomic ions are named in the same manner as binary ionic compounds. The name of the cation is given first, followed by the name of the anion. For example, the two compounds formed with silver by the nitrate and nitrite anions are named silver nitrate,  $\text{AgNO}_3$ , and silver nitrite,  $\text{AgNO}_2$ , respectively. When more than one polyatomic ion is present in a compound, the formula for the entire ion is surrounded by parentheses. This is illustrated on page 204 for aluminum sulfate,  $\text{Al}_2(\text{SO}_4)_3$ . The formula indicates that an aluminum sulfate formula unit has 2 aluminum cations and 3 sulfate anions.

**SAMPLE PROBLEM 7-3****Write the formula for tin(IV) sulfate.****SOLUTION***Write the symbols for the ions side by side. Write the cation first.**Cross over the charges to give subscripts. Add parentheses around the polyatomic ion if necessary.**Check the subscripts and write the formula.*

The total positive charge is  $2 \times 4+ = 8+$ . The total negative charge is  $4 \times 2- = 8-$ . The charges are equal. The largest common factor of the subscripts is 2, so the smallest whole-number ratio of ions in the compound is 1:2. The correct formula is therefore  $\text{Sn}(\text{SO}_4)_2$ .

**PRACTICE****1. Write formulas for the following ionic compounds:**

- |                      |                          |
|----------------------|--------------------------|
| a. sodium iodide     | e. copper(II) sulfate    |
| b. calcium chloride  | f. sodium carbonate      |
| c. potassium sulfide | g. calcium nitrite       |
| d. lithium nitrate   | h. potassium perchlorate |

**2. Give the names for the following compounds:**

- $\text{Ag}_2\text{O}$
- $\text{Ca}(\text{OH})_2$
- $\text{KClO}_3$
- $\text{NH}_4\text{OH}$
- $\text{FeCrO}_4$
- $\text{KClO}$

*Answer*

- |                         |                               |
|-------------------------|-------------------------------|
| 1. a. $\text{NaI}$      | e. $\text{CuSO}_4$            |
| b. $\text{CaCl}_2$      | f. $\text{Na}_2\text{CO}_3$   |
| c. $\text{K}_2\text{S}$ | g. $\text{Ca}(\text{NO}_2)_2$ |
| d. $\text{LiNO}_3$      | h. $\text{KClO}_4$            |
- silver oxide
    - calcium hydroxide
    - potassium chlorate
    - ammonium hydroxide
    - iron(II) chromate
    - potassium hypochlorite

## Naming Binary Molecular Compounds

Unlike ionic compounds, molecular compounds are composed of individual covalently bonded units, or molecules. Chemists use two nomenclature systems to name binary molecules. The newer system is the Stock system for naming molecular compounds, which requires an understanding of oxidation numbers. This system will be discussed in Section 7-2.

The old system of naming molecular compounds is based on the use of prefixes. For example, the molecular compound  $\text{CCl}_4$  is named carbon tetrachloride. The prefix *tetra-* indicates that four chloride atoms are present in a single molecule of the compound. The two oxides of carbon,  $\text{CO}$  and  $\text{CO}_2$ , are named carbon *monoxide* and carbon *dioxide*, respectively.

**TABLE 7-3 Numerical Prefixes**

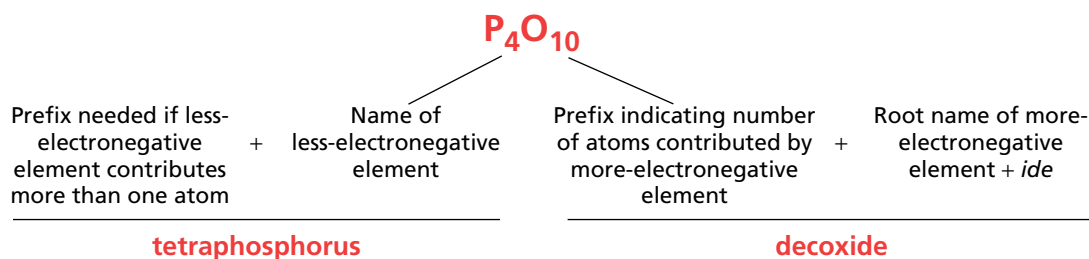
Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

In these names the prefix *mon-* indicates one oxygen atom and the prefix *di-* indicates two oxygen atoms. The prefixes used to specify the number of atoms, or sometimes the number of *groups* of atoms, in a molecule are listed in Table 7-3.

The rules for the prefix system of nomenclature of binary molecular compounds are as follows.

1. The less-electronegative element is given first. It is given a prefix only if it contributes more than one atom to a molecule of the compound.
2. The second element is named by combining (a) a prefix indicating the number of atoms contributed by the element, (b) the root of the name of the second element, and (c) the ending *-ide*. With few exceptions, the ending *-ide* indicates that a compound contains only two elements.
3. The *o* or *a* at the end of a prefix is usually dropped when the word following the prefix begins with another vowel, e.g., monoxide or pentoxide.

The prefix system is illustrated below.



Because the less-electronegative element is written first, oxygen and the halogens are usually given second, as in carbon tetrachloride and the carbon oxides. In general, the order of nonmetals in binary compound names and formulas is C, P, N, H, S, I, Br, Cl, O, F.

**TABLE 7-4 Binary Compounds of Nitrogen and Oxygen**

Formula	Prefix-system name
N <sub>2</sub> O	dinitrogen monoxide
NO	nitrogen monoxide
NO <sub>2</sub>	nitrogen dioxide
N <sub>2</sub> O <sub>3</sub>	dinitrogen trioxide
N <sub>2</sub> O <sub>4</sub>	dinitrogen tetroxide
N <sub>2</sub> O <sub>5</sub>	dinitrogen pentoxide

The prefix system is illustrated further in Table 7-4, which lists the names of the six oxides of nitrogen. Note the application of rule 1, for example, in the name *nitrogen dioxide* for NO<sub>2</sub>. No prefix is needed with *nitrogen* because only one atom of nitrogen, the less-electronegative element, is present in a molecule of NO<sub>2</sub>. On the other hand, the prefix *di-* in *dioxide* is needed according to rule 2 to indicate the presence of two atoms of the more-electronegative element, oxygen. Take a moment to review the prefixes in the other names in Table 7-4.

**SAMPLE PROBLEM 7-4**

- Give the name for As<sub>2</sub>O<sub>5</sub>.
- Write the formula for oxygen difluoride.

**SOLUTION**

- A molecule of the compound contains two arsenic atoms, so the first word in the name is “diarsenic.” The five oxygen atoms are indicated by adding the prefix *pent-* to the word “oxide.” The complete name is diarsenic pentoxide.
- The first symbol in the formula is that for oxygen. Oxygen is first in the name because it is less electronegative than fluorine. Since there is no prefix, there must be only one oxygen atom. The prefix *di-* in *difluoride* shows that there are two fluorine atoms in the molecule. The formula is OF<sub>2</sub>.

**PRACTICE**

- Name the following binary molecular compounds:
  - SO<sub>3</sub>
  - ICl<sub>3</sub>
  - PBr<sub>5</sub>
- Write formulas for the following compounds:
  - carbon tetraiodide
  - phosphorus trichloride
  - dinitrogen trioxide

*Answer*

- sulfur trioxide
  - iodine trichloride
  - phosphorus pentabromide
- Cl<sub>4</sub>
  - PCl<sub>3</sub>
  - N<sub>2</sub>O<sub>3</sub>

## Covalent-Network Compounds

As you read in Chapter 6, some covalent compounds do not consist of individual molecules. Instead, each atom is joined to all its neighbors in a covalently bonded, three-dimensional network. There are no distinct units in these compounds, just as there are no such units in ionic compounds. The subscripts in a formula for a covalent-network compound indicate the smallest whole-number ratio of the atoms in the compound. Naming such compounds is similar to naming molecular compounds. Some common examples are given below.

$\text{SiC}$	$\text{SiO}_2$	$\text{Si}_3\text{N}_4$
silicon carbide	silicon dioxide	trisilicon tetranitride

## Acids and Salts

An *acid* is a distinct type of molecular compound about which you will read much more in Chapter 15. Most acids used in the laboratory can be classified as either binary acids or oxyacids. *Binary acids* are acids that consist of two elements, usually hydrogen and one of the halogens—fluorine, chlorine, bromine, iodine. *Oxyacids* are acids that contain hydrogen, oxygen, and a third element (usually a nonmetal).

Acids were first recognized as a specific class of compounds based on their properties in solutions of water. Consequently, in chemical nomenclature, the term *acid* usually refers to a solution in water of one of these special compounds rather than to the compound itself. For example, *hydrochloric acid* refers to a water solution of the molecular compound hydrogen chloride, HCl. Some common binary and oxyacids are listed in Table 7-5.

Many polyatomic ions are produced by the loss of hydrogen ions from oxyacids. A few examples of the relationship between oxyacids and oxyanions are shown below.

sulfuric acid	$\text{H}_2\text{SO}_4$	sulfate	$\text{SO}_4^{2-}$
nitric acid	$\text{HNO}_3$	nitrate	$\text{NO}_3^-$
phosphoric acid	$\text{H}_3\text{PO}_4$	phosphate	$\text{PO}_4^{3-}$

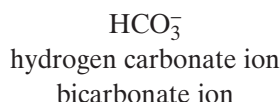
**TABLE 7-5** Common Binary Acids and Oxyacids

HF	hydrofluoric acid	$\text{HNO}_2$	nitrous acid	HClO	hypochlorous acid
HCl	hydrochloric acid	$\text{HNO}_3$	nitric acid	$\text{HClO}_2$	chlorous acid
HBr	hydrobromic acid	$\text{H}_2\text{SO}_3$	sulfurous acid	$\text{HClO}_3$	chloric acid
HI	hydriodic acid	$\text{H}_2\text{SO}_4$	sulfuric acid	$\text{HClO}_4$	perchloric acid
$\text{H}_3\text{PO}_4$	phosphoric acid	$\text{CH}_3\text{COOH}$	acetic acid	$\text{H}_2\text{CO}_3$	carbonic acid



**FIGURE 7-2** Some common laboratory acids. Acids should always be handled with care and according to instructions. They can burn the skin, and they burn holes in clothing.

An ionic compound composed of a cation and the anion from an acid is often referred to as a **salt**. Table salt,  $\text{NaCl}$ , contains the anion from hydrochloric acid. Calcium sulfate,  $\text{CaSO}_4$ , is a salt containing an anion from sulfuric acid. Some salts contain anions in which one or more hydrogen atoms from the acid are retained. Such anions are named by adding the word *hydrogen* or the prefix *bi-* to the anion name. The best known such anion comes from carbonic acid,  $\text{H}_2\text{CO}_3$ .



## SECTION REVIEW

- What is the significance of a chemical formula?
- Write formulas for the compounds formed between the following:
  - aluminum and bromine
  - sodium and oxygen
  - magnesium and iodine
  - $\text{Pb}^{2+}$  and  $\text{O}^{2-}$
  - $\text{Sn}^{2+}$  and  $\text{I}^-$
  - $\text{Fe}^{3+}$  and  $\text{S}^{2-}$
  - $\text{Cu}^{2+}$  and  $\text{NO}_3^-$
  - $\text{NH}_4^+$  and  $\text{SO}_4^{2-}$
- Name the following compounds using the Stock system:
 

a. $\text{NaI}$	c. $\text{CaO}$	e. $\text{CuBr}$
b. $\text{MgS}$	d. $\text{K}_2\text{S}$	f. $\text{FeCl}_2$
- Write formulas for each of the following compounds:
 

a. barium sulfide	g. disulfur dichloride
b. sodium hydroxide	h. carbon diselenide
c. lead(II) nitrate	i. acetic acid
d. potassium permanganate	j. chloric acid
e. iron(II) sulfate	k. sulfurous acid
f. diphosphorus trioxide	l. phosphoric acid



## OBJECTIVES

- List the rules for assigning oxidation numbers.
- Give the oxidation number for each element in the formula of a chemical compound.
- Name binary molecular compounds using oxidation numbers and the Stock system.

# Oxidation Numbers

The charges on the ions composing an ionic compound reflect the electron distribution of the compound. *In order to indicate the general distribution of electrons among the bonded atoms in a molecular compound or a polyatomic ion, **oxidation numbers**, also called **oxidation states**, are assigned to the atoms composing the compound or ion.* Unlike ionic charges, oxidation numbers do not have an exact physical meaning. In fact, in some cases they are quite arbitrary. However, oxidation numbers are useful in naming compounds, in writing formulas, and in balancing chemical equations. And, as will be discussed in Chapter 19, they are helpful in studying certain types of chemical reactions.

## Assigning Oxidation Numbers

As a general rule in assigning oxidation numbers, shared electrons are assumed to belong to the more-electronegative atom in each bond. More specific rules for determining oxidation numbers are provided by the following guidelines.

1. The atoms in a pure element have an oxidation number of zero. For example, the atoms in pure sodium, Na, oxygen, O<sub>2</sub>, phosphorus, P<sub>4</sub>, and sulfur, S<sub>8</sub>, all have oxidation numbers of zero.
2. The more-electronegative element in a binary molecular compound is assigned the number equal to the negative charge it would have as an anion. The less-electronegative atom is assigned the number equal to the positive charge it would have as a cation.
3. Fluorine has an oxidation number of  $-1$  in all of its compounds because it is the most electronegative element.
4. Oxygen has an oxidation number of  $-2$  in almost all compounds. Exceptions include when it is in peroxides, such as H<sub>2</sub>O<sub>2</sub>, in which its oxidation number is  $-1$ , and when it is in compounds with halogens, such as OF<sub>2</sub>, in which its oxidation number is  $+2$ .
5. Hydrogen has an oxidation number of  $+1$  in all compounds containing elements that are more-electronegative than it; it has an oxidation number of  $-1$  in compounds with metals.
6. The algebraic sum of the oxidation numbers of all atoms in a neutral compound is equal to zero.
7. The algebraic sum of the oxidation numbers of all atoms in a polyatomic ion is equal to the charge of the ion.
8. Although rules 1 through 7 apply to covalently bonded atoms, oxidation numbers can also be assigned to atoms in ionic compounds.

A monatomic ion has an oxidation number equal to the charge of the ion. For example, the ions  $\text{Na}^+$ ,  $\text{Ca}^{2+}$ , and  $\text{Cl}^-$  have oxidation numbers of +1, +2, and -1, respectively.

Let's examine the assignment of oxidation numbers to the atoms in two molecular compounds, hydrogen fluoride, HF, and water,  $\text{H}_2\text{O}$ . In HF the bond is polar, with a partial negative charge on the fluorine atom and a partial positive charge on the hydrogen atom. If HF were an ionic compound in which an electron was fully transferred to the fluorine atom, H would have a 1+ charge and F would have a 1- charge. Thus, the oxidation numbers of H and F in hydrogen fluoride are +1 and -1, respectively. In a water molecule, the oxygen atom is more electronegative than the hydrogen atoms. If  $\text{H}_2\text{O}$  were an ionic compound, the oxygen atom would have a charge of 2- and the hydrogen atoms would each have a charge of 1+. The oxidation numbers of H and O in water are therefore +1 and -2, respectively.

Because the sum of the oxidation numbers of the atoms in a compound must satisfy rule 6 or 7 of the guidelines on page 216, it is often possible to assign oxidation numbers when they are not known. This is illustrated in Sample Problem 7-5.

### SAMPLE PROBLEM 7-5

**Assign oxidation numbers to each atom in the following compounds or ions:**

- a.  $\text{UF}_6$
- b.  $\text{H}_2\text{SO}_4$
- c.  $\text{ClO}_3^-$

#### SOLUTION

- a. Start by placing known oxidation numbers above the appropriate elements. From the guidelines, we know that fluorine always has an oxidation number of -1.



Multiply known oxidation numbers by the appropriate number of atoms and place the totals underneath the corresponding elements. There are six fluorine atoms,  $6 \times -1 = -6$ .



The compound  $\text{UF}_6$  is molecular. According to the guidelines, the sum of the oxidation numbers must equal zero. The total of positive oxidation numbers is therefore +6.



Divide the total calculated oxidation number by the appropriate number of atoms. There is only one uranium atom in the molecule, so it must have an oxidation number of +6.

+6 -1

UF<sub>6</sub>

+6 -6

- b. Oxygen and sulfur are each more electronegative than hydrogen, so hydrogen has an oxidation number of +1. Oxygen is not combined with a halogen, nor is H<sub>2</sub>SO<sub>4</sub> a peroxide. Therefore, the oxidation number of oxygen is -2. Place these known oxidation numbers above the appropriate symbols. Place the total of the oxidation numbers underneath.

+1 -2

H<sub>2</sub>SO<sub>4</sub>

+2 -8

The sum of the oxidation numbers must equal zero, and there is only one sulfur atom in each molecule of H<sub>2</sub>SO<sub>4</sub>. Each sulfur atom therefore must have an oxidation number of (+2) + (-8) = +6.

- c. To assign oxidation numbers to the elements in ClO<sub>3</sub><sup>-</sup>, proceed as in parts (a) and (b). Remember, however, that the total of the oxidation numbers should equal the overall charge of the anion, 1-. The oxidation number of a single oxygen atom in the ion is -2. The total oxidation number due to the three oxygen atoms is -6. For the chlorate ion to have a 1- charge, chlorine must be assigned an oxidation number of +5.

+5 -2

ClO<sub>3</sub><sup>-</sup>

+5 -6

## PRACTICE

1. Assign oxidation numbers to each atom in the following compounds or ions:

a. HCl    e. HNO<sub>3</sub>    h. HClO<sub>3</sub>  
b. CF<sub>4</sub>    f. KH    i. N<sub>2</sub>O<sub>5</sub>  
c. PCl<sub>3</sub>    g. P<sub>4</sub>O<sub>10</sub>    j. GeCl<sub>2</sub>  
d. SO<sub>2</sub>

*Answer*

a. +1, -1    e. +1, +5, -2    h. +1, +5, -2  
b. +4, -1    f. +1, -1    i. +5, -2  
c. +3, -1    g. +5, -2    j. +2, -1  
d. +4, -2

## Using Oxidation Numbers for Formulas and Names

As shown in Table 7-6, many nonmetals can have more than one oxidation number. (A more extensive list of oxidation numbers is given in Appendix Table A-15.) These numbers can sometimes be used in the same manner as ionic charges to determine formulas. Suppose, for example, you want to know the formula of a binary compound formed between sulfur and oxygen. From the common +4 and +6 oxidation states of sulfur, you could expect that sulfur might form SO<sub>2</sub> or SO<sub>3</sub>. Both are known compounds. Of course, a formula must represent facts. Oxidation numbers alone cannot be used to predict the existence of a compound.

**TABLE 7-6** Common Oxidation Numbers of Some Nonmetals That Have Variable Oxidation States\*

Group 14	carbon	-4, +2, +4
Group 15	nitrogen	-3, +3, +5
	phosphorus	-3, +3, +5
Group 16	sulfur	-2, +4, +6
Group 17	chlorine	-1, +1, +3, +5, +7
	bromine	-1, +1, +3, +5, +7
	iodine	-1, +1, +3, +5, +7

\*In addition to the values shown, atoms of each element in its pure state are assigned an oxidation number of zero.

In Section 7-1 we introduced the use of Roman numerals to denote ionic charges in the Stock system of naming ionic compounds. The Stock system is actually based on oxidation numbers, and it can be used as an alternative to the prefix system for naming binary molecular compounds. In the prefix system, for example,  $\text{SO}_2$  and  $\text{SO}_3$  are named sulfur dioxide and sulfur trioxide, respectively. Their names according to the Stock system are sulfur(IV) oxide and sulfur(VI) oxide. The international body that governs nomenclature has endorsed the Stock system, which is more practical for complicated compounds. Prefix-based names and Stock-system names are still used interchangeably for many simple compounds, however. A few additional examples of names in both systems are given below.

	Prefix system	Stock system
$\text{PCl}_3$	phosphorus trichloride	phosphorus(III) chloride
$\text{PCl}_5$	phosphorus pentachloride	phosphorus(V) chloride
$\text{N}_2\text{O}$	dinitrogen monoxide	nitrogen(I) oxide
$\text{NO}$	nitrogen monoxide	nitrogen(II) oxide
$\text{PbO}_2$	lead dioxide	lead(IV) oxide
$\text{Mo}_2\text{O}_3$	dimolybdenum trioxide	molybdenum(III) oxide

## SECTION REVIEW

- Assign oxidation numbers to each atom in the following compounds or ions:
  - $\text{HF}$
  - $\text{Cl}_4$
  - $\text{H}_2\text{O}$
  - $\text{PI}_3$
  - $\text{CS}_2$
  - $\text{Na}_2\text{O}_2$
  - $\text{H}_2\text{CO}_3$
  - $\text{NO}_2^-$
- Name each of the following binary molecular compounds according to the Stock system:
  - $\text{Cl}_4$
  - $\text{SO}_3$
  - $\text{As}_2\text{S}_3$
  - $\text{NCl}_3$
  - $\text{SO}_4^{2-}$
  - $\text{ClO}_2^-$
  - $\text{IO}_3^-$



## Chemistry and Art

From "Chemistry and the Plastic and Graphic Arts" by Jonathan E. Ericson  
in *The Central Science, Essays on the Uses of Chemistry*

**T**he preservation of art is dependent on the control of the environment of the piece. Modern museums are air conditioned, maintaining their temperature between 68°–72°F and a 50–65% relative humidity. These controls provide the conditions under which most works of art are stable. Occasionally, the relative humidity will have to be increased or decreased depending on the stability of an individual piece. Here, art conservation has borrowed techniques from physical chemistry to determine proper conditions experimentally.

The lighting of a work of art is also a critical part of its environment. Fluorescent light and sunlight contain a lot of ultraviolet. The exposure of art to this light can cause it to fade tremendously. Papers, textiles, and organic dyes are the most sensitive to ultraviolet fading. Here, the polymer research chemist has developed special acrylic plastics like Plexiglas® UF-3, a Rohm and Haas product, which filters out the ultraviolet light.

The relative humidity, temperature and lighting are rather easy to control in a large museum. However, the agents of attack which deteriorate the appearance of the art piece are now not always so easily identified. For paintings, the accumulation of dirt and grime, dis-



*In many museums, the temperature and relative humidity of the surroundings are recorded on the same graph by a hygrothermograph, as shown here.*

coloring of the protective varnish layer, or the tension and distortion of canvas or wood supports are destructive to the piece and distorting to us when viewing it. In the past, when cleaning a painting, people either used stronger cleaning methods like sandpaper, or, if smart, gave up. Modern chemistry has developed ways and means of using whole families of safer cleaning agents like acetone, alcohol and other organic solvents. After careful examination and evaluation, the painting conservator can remove the old varnish, reline or bond a new supporting canvas with natural wax resin or synthetic resin, spray on an isolating layer of synthetic resin, paint in pigment losses, and seal the painting with a final protective layer. All these steps use products which have been developed in the chemical laboratory.

Finally, the agents of attack are not always so subtle. Sometimes works of art fall and break, or are torn, cut, or burned. For each case and for each object, a particular conservation method is used. Almost always, the products used or the treatment itself are direct contributions of a chemist. Without chemistry, art conservation, as we know it, would be very primitive, indeed.

### Reading for Meaning

What kind of molecules are acrylic plastics composed of?

### Read Further

When light interacts with paper, it can cause a *photochemical reaction*. Research photochemical reactions, and explain why a newspaper left in the sunlight for a long period of time turns yellow.

# Using Chemical Formulas

## SECTION 7-3

### OBJECTIVES

- Calculate the formula mass or molar mass of any given compound.
- Use molar mass to convert between mass in grams and amount in moles of a chemical compound.
- Calculate the number of molecules, formula units, or ions in a given molar amount of a chemical compound.
- Calculate the percentage composition of a given chemical compound.

### Formula Masses

In Chapter 3 we saw that hydrogen atoms have an average atomic mass of 1.007 94 amu and that oxygen atoms have an average atomic mass of 15.9994 amu. Like individual atoms, a molecule, a formula unit, or an ion has a characteristic average mass. For example, we know from the chemical formula  $\text{H}_2\text{O}$  that a single water molecule is composed of exactly two hydrogen atoms and one oxygen atom. The mass of a water molecule is found by summing the masses of the three atoms in the molecule. (In the calculation, the average atomic masses have been rounded to two decimal places.)


average atomic mass of H: 1.01 amu  
average atomic mass of O: 16.00 amu

$$2 \text{ H-atoms} \times \frac{1.01 \text{ amu}}{\text{H-atom}} = 2.02 \text{ amu}$$

$$1 \text{ O-atom} \times \frac{16.00 \text{ amu}}{\text{O-atom}} = 16.00 \text{ amu}$$

average mass of  $\text{H}_2\text{O}$  molecule = 18.02 amu

The mass of a water molecule can be correctly referred to as a *molecular mass*. The mass of one NaCl formula unit, on the other hand, is not a molecular mass because NaCl is an ionic compound. The mass of any unit represented by a chemical formula, whether the unit is a molecule, a formula unit, or an ion, is known as the formula mass. *The formula mass of any molecule, formula unit, or ion is the sum of the average atomic masses of all the atoms represented in its formula.*

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The procedure illustrated for calculating the formula mass of a water molecule can be used to calculate the mass of any unit represented by a chemical formula. In each of the problems that follow, the atomic masses from the periodic table on pages 130–131 have been rounded to two decimal places.

### SAMPLE PROBLEM 7-6

**Find the formula mass of potassium chlorate,  $\text{KClO}_3$ .**

#### SOLUTION

The mass of a formula unit of  $\text{KClO}_3$  is found by summing the masses of one K atom, one Cl atom, and three O atoms. The required atomic masses can be found in the periodic table on pages 130–131. In the calculation, each atomic mass has been rounded to two decimal places.

$$1 \text{ K-atom} \times \frac{39.10 \text{ amu}}{\text{K-atom}} = 39.10 \text{ amu}$$

$$1 \text{ Cl-atom} \times \frac{35.45 \text{ amu}}{\text{Cl-atom}} = 35.45 \text{ amu}$$

$$3 \text{ O-atoms} \times \frac{16.00 \text{ amu}}{\text{O-atom}} = 48.00 \text{ amu}$$

$$\text{formula mass of } \text{KClO}_3 = 122.55 \text{ amu}$$

#### PRACTICE

1. Find the formula mass of each of the following:

- a.  $\text{H}_2\text{SO}_4$
- b.  $\text{Ca}(\text{NO}_3)_2$
- c.  $\text{PO}_4^{3-}$
- d.  $\text{MgCl}_2$

*Answer*

- a. 98.08 amu
- b. 164.10 amu
- c. 94.97 amu
- d. 95.21 amu

## Molar Masses

In Chapter 3 you learned that the molar mass of a substance is equal to the mass in grams of one mole, or approximately  $6.022 \times 10^{23}$  particles, of the substance. For example, the molar mass of pure calcium, Ca, is 40.08 g/mol because one mole of calcium atoms has a mass of 40.08 g.

The molar mass of a compound is calculated by summing the masses of the elements present in a mole of the molecules or formula units that make up the compound. For example, one mole of water molecules contains exactly two moles of H atoms and one mole of O atoms. Rounded to two decimal places, a mole of hydrogen atoms has a mass

of 1.01 g, and a mole of oxygen atoms has a mass of 16.00 g. The molar mass of water is calculated as follows.

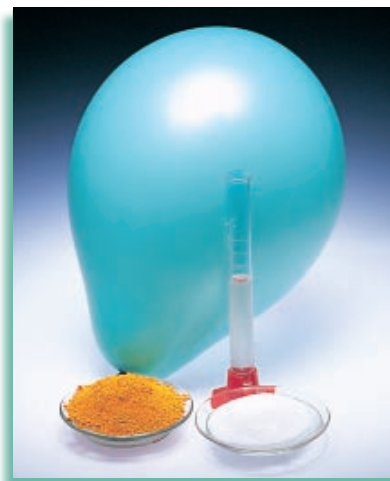
$$2 \text{ mol H} \times \frac{1.01 \text{ g H}}{\text{mol H}} = 2.02 \text{ g H}$$

$$1 \text{ mol O} \times \frac{16.00 \text{ g O}}{\text{mol O}} = 16.00 \text{ g O}$$

$$\text{molar mass of H}_2\text{O} = 18.02 \text{ g/mol}$$

Figure 7-3 shows a mole of water as well as a mole of several other substances.

You may have noticed that *a compound's molar mass is numerically equal to its formula mass*. For instance, in Sample Problem 7-6 the formula mass of  $\text{KClO}_3$  was found to be 122.55 amu. Therefore, because molar mass is numerically equal to formula mass, we know that the molar mass of  $\text{KClO}_3$  is 122.55 g/mol.



**FIGURE 7-3** Every compound has a characteristic molar mass. Shown here are one mole each of nitrogen (in balloon), water (in graduated cylinder), cadmium sulfide,  $\text{CdS}$  (yellow substance), and sodium chloride,  $\text{NaCl}$  (white substance).

### SAMPLE PROBLEM 7-7

**What is the molar mass of barium nitrate,  $\text{Ba}(\text{NO}_3)_2$ ?**

#### SOLUTION

One mole of barium nitrate contains exactly one mole of  $\text{Ba}^{2+}$  ions and two moles of  $\text{NO}_3^-$  ions. The two moles of  $\text{NO}_3^-$  ions contain two moles of N atoms and six moles of O atoms. Therefore, the molar mass of  $\text{Ba}(\text{NO}_3)_2$  is calculated as follows.

$$1 \text{ mol Ba} \times \frac{137.33 \text{ g Ba}}{\text{mol Ba}} = 137.33 \text{ g Ba}$$

$$2 \text{ mol N} \times \frac{14.01 \text{ g N}}{\text{mol N}} = 28.02 \text{ g N}$$

$$6 \text{ mol O} \times \frac{16.00 \text{ g O}}{\text{mol O}} = 96.00 \text{ g O}$$

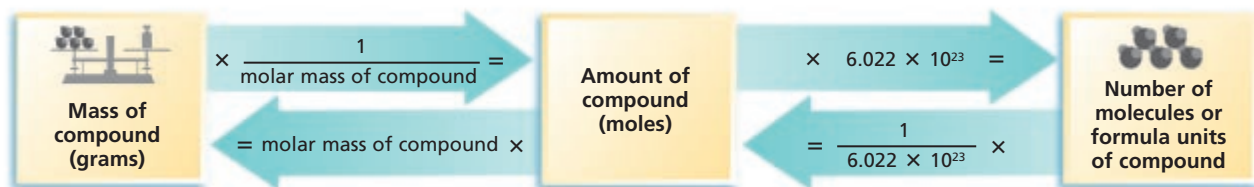
$$\text{molar mass of Ba}(\text{NO}_3)_2 = 261.35 \text{ g/mol}$$

#### PRACTICE

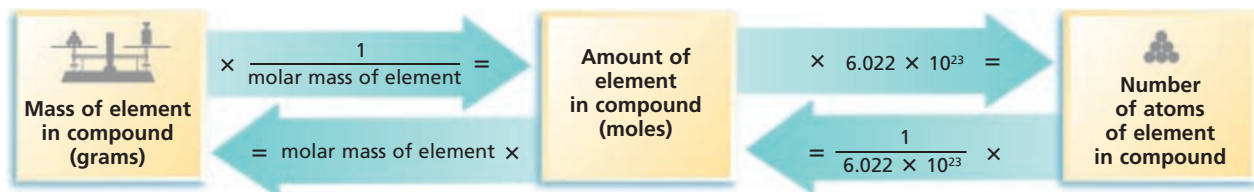
- How many moles of atoms of each element are there in one mole of the following compounds?
  - $\text{Al}_2\text{S}_3$
  - $\text{NaNO}_3$
  - $\text{Ba}(\text{OH})_2$
- Find the molar mass of each of the compounds listed in item 1.

*Answer*

- 2 mol Al, 3 mol S
  - 1 mol Na, 1 mol N, 3 mol O
  - 1 mol Ba, 2 mol O, 2 mol H
- 150.17 g/mol
  - 85.00 g/mol
  - 171.35 g/mol



(a)



(b)

**FIGURE 7-4** (a) The diagram shows the relationships between mass in grams, amount in moles, and number of molecules or atoms for a given compound. (b) Similar relationships exist for an element within a compound.

## Molar Mass as a Conversion Factor

The molar mass of a compound can be used as a conversion factor to relate an amount in moles to a mass in grams for a given substance. Recall that molar mass usually has the units of grams per mole. To convert a known amount of a compound in moles to a mass in grams, multiply the amount in moles by the molar mass.

$$\text{amount in moles} \times \text{molar mass (g/mol)} = \text{mass in grams}$$

Conversions of this type for elements and compounds are summarized above in Figure 7-4.

### SAMPLE PROBLEM 7-8

What is the mass in grams of 2.50 mol of oxygen gas?

#### SOLUTION

##### 1 ANALYZE

**Given:** 2.50 mol O<sub>2</sub>

**Unknown:** mass of O<sub>2</sub> in grams

##### 2 PLAN

moles O<sub>2</sub> → grams O<sub>2</sub>

To convert amount of O<sub>2</sub> in moles to mass of O<sub>2</sub> in grams, multiply by the molar mass of O<sub>2</sub>.

$$\text{amount of O}_2 \text{ (mol)} \times \text{molar mass of O}_2 \text{ (g/mol)} = \text{mass of O}_2 \text{ (g)}$$

- 3 COMPUTE** First the molar mass of O<sub>2</sub> must be calculated.

$$2 \cancel{\text{mol O}} \times \frac{16.00 \text{ g O}}{\cancel{\text{mol O}}} = 32.00 \text{ g (mass of one mole of O}_2\text{)}$$

The molar mass of O<sub>2</sub> is therefore 32.00 g/mol. Now do the calculation shown in step 2.

$$2.50 \cancel{\text{mol O}_2} \times \frac{32.00 \text{ g O}_2}{\cancel{\text{mol O}_2}} = 80.0 \text{ g O}_2$$

- 4 EVALUATE** The answer is correctly given to three significant figures and is close to an estimated value of 75 g (2.50 mol × 30 g/mol).

To convert a known mass of a compound in grams to an amount in moles, the mass must be divided by the molar mass. Or you can invert the molar mass and multiply so that units are easily canceled.

$$\text{mass in grams} \times \frac{1}{\text{molar mass (g/mol)}} = \text{amount in moles}$$

### SAMPLE PROBLEM 7-9

Ibuprofen, C<sub>13</sub>H<sub>18</sub>O<sub>2</sub>, is the active ingredient in many nonprescription pain relievers. Its molar mass is 206.29 g/mol.

- If the tablets in a bottle contain a total of 33 g of ibuprofen, how many moles of ibuprofen are in the bottle?
- How many molecules of ibuprofen are in the bottle?
- What is the total mass in grams of carbon in 33 g of ibuprofen?

### SOLUTION

**1 ANALYZE**

**Given:** 33 g of C<sub>13</sub>H<sub>18</sub>O<sub>2</sub>, molar mass 206.29 g/mol

**Unknown:** a. moles C<sub>13</sub>H<sub>18</sub>O<sub>2</sub>  
b. molecules C<sub>13</sub>H<sub>18</sub>O<sub>2</sub>  
c. total mass of C

**2 PLAN**

- a. grams → moles

To convert mass of ibuprofen in grams to amount of ibuprofen in moles, multiply by the inverted molar mass of C<sub>13</sub>H<sub>18</sub>O<sub>2</sub>.

$$\text{g C}_{13}\text{H}_{18}\text{O}_2 \times \frac{1 \text{ mol C}_{13}\text{H}_{18}\text{O}_2}{206.29 \text{ g C}_{13}\text{H}_{18}\text{O}_2} = \text{mol C}_{13}\text{H}_{18}\text{O}_2$$

- b. moles → molecules

To find the number of molecules of ibuprofen, multiply amount of C<sub>13</sub>H<sub>18</sub>O<sub>2</sub> in moles by Avogadro's number.

$$\text{mol C}_{13}\text{H}_{18}\text{O}_2 \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = \text{molecules C}_{13}\text{H}_{18}\text{O}_2$$

c. moles  $\text{C}_{13}\text{H}_{18}\text{O}_2 \longrightarrow$  moles C  $\longrightarrow$  grams C

To find the mass of carbon present in the ibuprofen, the two conversion factors needed are the amount of carbon in moles per mole of  $\text{C}_{13}\text{H}_{18}\text{O}_2$  and the molar mass of carbon.

$$\text{mol C}_{13}\text{H}_{18}\text{O}_2 \times \frac{13 \text{ mol C}}{\text{mol C}_{13}\text{H}_{18}\text{O}_2} \times \frac{12.01 \text{ g C}}{\text{mol C}} = \text{g C}$$

### 3 COMPUTE

a.  $33 \text{ g C}_{13}\text{H}_{18}\text{O}_2 \times \frac{1 \text{ mol C}_{13}\text{H}_{18}\text{O}_2}{206.29 \text{ g C}_{13}\text{H}_{18}\text{O}_2} = 0.16 \text{ mol C}_{13}\text{H}_{18}\text{O}_2$

b.  $0.16 \text{ mol C}_{13}\text{H}_{18}\text{O}_2 \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 9.6 \times 10^{22} \text{ molecules C}_{13}\text{H}_{18}\text{O}_2$

c.  $0.16 \text{ mol C}_{13}\text{H}_{18}\text{O}_2 \times \frac{13 \text{ mol C}}{\text{mol C}_{13}\text{H}_{18}\text{O}_2} \times \frac{12.01 \text{ g C}}{\text{mol C}} = 25 \text{ g C}$

The bottle contains 0.16 mol of ibuprofen, which is  $9.6 \times 10^{22}$  molecules of ibuprofen. The sample of ibuprofen contains 25 g of carbon.

### 4 EVALUATE

Checking each step shows that the arithmetic is correct, significant figures have been used correctly, and units have canceled as desired.

### PRACTICE

1. How many moles of compound are there in the following?

- a. 6.60 g  $(\text{NH}_4)_2\text{SO}_4$   
b. 4.5 kg  $\text{Ca}(\text{OH})_2$

*Answer*

1. a. 0.0500 mol  
b. 61 mol

2. How many molecules are there in the following?

- a. 25.0 g  $\text{H}_2\text{SO}_4$   
b. 125 g of sugar,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

2. a.  $1.53 \times 10^{23}$  molecules  
b.  $2.20 \times 10^{23}$  molecules

3. What is the mass in grams of 6.25 mol of copper(II) nitrate?

3. 1170 g

## Percentage Composition

It is often useful to know the percentage by mass of a particular element in a chemical compound. For example, suppose the compound potassium chlorate,  $\text{KClO}_3$ , were to be used as a source of oxygen. It would be helpful to know the percentage of oxygen in the compound. To find the mass percentage of an element in a compound, one can divide the mass of the element in a sample of the compound by the total mass of the sample, then multiply this value by 100.

$$\frac{\text{mass of element in sample of compound}}{\text{mass of sample of compound}} \times 100 = \% \text{ element in compound}$$

The mass percentage of an element in a compound is the same regardless of the sample's size. Therefore, a simpler way to calculate the percentage of an element in a compound is to determine how many grams of the element are present in one mole of the compound. Then divide this value by the molar mass of the compound and multiply by 100.

$$\frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100 = \begin{matrix} \% \text{ element in} \\ \text{compound} \end{matrix}$$

The percentage by mass of each element in a compound is known as the **percentage composition** of the compound.

### SAMPLE PROBLEM 7-10

Find the percentage composition of copper(I) sulfide,  $\text{Cu}_2\text{S}$ .

#### SOLUTION

##### 1 ANALYZE

**Given:** formula,  $\text{Cu}_2\text{S}$

**Unknown:** percentage composition of  $\text{Cu}_2\text{S}$

##### 2 PLAN

formula  $\longrightarrow$  molar mass  $\longrightarrow$  mass percentage of each element

The molar mass of the compound must be found. Then the mass of each element present in one mole of the compound is used to calculate the mass percentage of each element.

##### 3 COMPUTE

$$2 \cancel{\text{mol Cu}} \times \frac{63.55 \text{ g Cu}}{\cancel{\text{mol Cu}}} = 127.1 \text{ g Cu}$$

$$1 \cancel{\text{mol S}} \times \frac{32.07 \text{ g S}}{\cancel{\text{mol S}}} = 32.07 \text{ g S}$$

$$\text{molar mass of Cu}_2\text{S} = 159.2 \text{ g}$$

$$\frac{127.1 \text{ g Cu}}{159.2 \text{ g Cu}_2\text{S}} \times 100 = 79.84\% \text{ Cu}$$

$$\frac{32.07 \text{ g S}}{159.2 \text{ g Cu}_2\text{S}} \times 100 = 20.14\% \text{ S}$$

##### 4 EVALUATE

A good check is to see if the results add up to about 100%. (Because of rounding, the total may not always be exactly 100%.)

### SAMPLE PROBLEM 7-11

As some salts crystallize from a water solution, they bind water molecules in their crystal structure. Sodium carbonate forms such a *hydrate*, in which 10 water molecules are present for every formula unit of sodium carbonate. Find the mass percentage of water in sodium carbonate decahydrate,  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ , which has a molar mass of 286.14 g/mol.



## SOLUTION

- 1 ANALYZE** **Given:** chemical formula,  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$   
molar mass of  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$   
**Unknown:** mass percentage of  $\text{H}_2\text{O}$
- 2 PLAN** chemical formula  $\longrightarrow$  mass  $\text{H}_2\text{O}$  per mole of  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$   $\longrightarrow$  % water
- The mass of water per mole of sodium carbonate decahydrate must first be found. This value is then divided by the mass of one mole of  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ .
- 3 COMPUTE** One mole of  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  contains 10 mol of  $\text{H}_2\text{O}$ . Recall from page 223 that the molar mass of  $\text{H}_2\text{O}$  is 18.02 g/mol. The mass of 10 mol of  $\text{H}_2\text{O}$  is calculated as follows.
- $$10 \text{ mol } \text{H}_2\text{O} \times \frac{18.02 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \text{H}_2\text{O}} = 180.2 \text{ g } \text{H}_2\text{O}$$
- $$\text{mass of } \text{H}_2\text{O} \text{ per mole of } \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = 180.2 \text{ g}$$
- The molar mass of  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  is 286.14 g/mol, so we know that 1 mol of the hydrate has a mass of 286.14g. The mass percentage of 10 mol of  $\text{H}_2\text{O}$  in 1 mol of  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  can now be calculated.
- $$\text{mass percentage of } \text{H}_2\text{O} \text{ in } \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = \frac{180.2 \text{ g } \text{H}_2\text{O}}{286.14 \text{ g } \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}} \times 100 = 62.98\% \text{ H}_2\text{O}$$
- 4 EVALUATE** Checking shows that the arithmetic is correct and that units cancel as desired.

## PRACTICE

- |   |                                  |
|---|----------------------------------|
| 1. Find the percentage compositions of the following:   | <i>Answer</i>                    |
| a. $\text{PbCl}_2$  | 1. a. 74.51% Pb, 25.49% Cl       |
| b. $\text{Ba}(\text{NO}_3)_2$   | b. 52.55% Ba, 10.72% N, 36.73% O |
| 2. Find the mass percentage of water in $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ .   | 2. 43.86% $\text{H}_2\text{O}$   |
| 3. Magnesium hydroxide is 54.87% oxygen by mass. How many grams of oxygen are in 175 g of the compound? How many moles of oxygen is this? | 3. 96.0 g O; 6.00 mol O          |

## SECTION REVIEW

- |   |   |
|---|---|
| 1. Determine both the formula mass and molar mass of ammonium carbonate, $(\text{NH}_4)_2\text{CO}_3$ . | 4. How many moles of molecules are there in 250 g of hydrogen nitrate, $\text{HNO}_3$ ?                         |
| 2. How many moles of atoms of each element are there in one mole of $(\text{NH}_4)_2\text{CO}_3$ ?      | 5. How many molecules of aspirin, $\text{C}_9\text{H}_8\text{O}_4$ , are there in a 100.0 mg tablet of aspirin? |
| 3. What is the mass in grams of 3.25 mol $\text{Fe}_2(\text{SO}_4)_3$ ?                                 | 6. Calculate the percentage composition of $(\text{NH}_4)_2\text{CO}_3$ .                                       |

# Determining Chemical Formulas

## SECTION 7-4

### OBJECTIVES

- Define *empirical formula*, and explain how the term applies to ionic and molecular compounds.
- Determine an empirical formula from either a percentage or a mass composition.
- Explain the relationship between the empirical formula and the molecular formula of a given compound.
- Determine a molecular formula from an empirical formula.

When a new substance is synthesized or is discovered, it is analyzed quantitatively to reveal its percentage composition. From this data, the empirical formula is then determined. An **empirical formula** consists of the symbols for the elements combined in a compound, with subscripts showing the smallest whole-number mole ratio of the different atoms in the compound. For an ionic compound, the formula unit is usually the compound's empirical formula. For a molecular compound, however, the empirical formula does not necessarily indicate the actual numbers of atoms present in each molecule. For example, the empirical formula of the gas diborane is  $\text{BH}_3$ , but the molecular formula is  $\text{B}_2\text{H}_6$ . In this case, the number of atoms given by the molecular formula corresponds to the empirical ratio multiplied by two.

### Calculation of Empirical Formulas

To determine a compound's empirical formula from its percentage composition, begin by converting percentage composition to a mass composition. Assume that you have a 100.0 g sample of the compound. Then calculate the amount of each element in the sample. For example, the percentage composition of diborane is 78.1% B and 21.9% H. Therefore, 100.0 g of diborane contains 78.1 g of B and 21.9 g of H.

Next, the mass composition of each element is converted to a composition in moles by dividing by the appropriate molar mass.

$$78.1 \text{ g B} \times \frac{1 \text{ mol B}}{10.81 \text{ g B}} = 7.22 \text{ mol B}$$

$$21.9 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 21.7 \text{ mol H}$$

These values give a mole ratio of 7.22 mol B to 21.7 mol H. However, this is not a ratio of smallest whole numbers. To find such a ratio, divide each number of moles by the smallest number in the existing ratio.

$$\frac{7.22 \text{ mol B}}{7.22} : \frac{21.7 \text{ mol H}}{7.22} = 1 \text{ mol B} : 3.01 \text{ mol H}$$

Because of rounding or experimental error, a compound's mole ratio sometimes consists of numbers close to whole numbers instead of exact whole numbers. In this case, the differences from whole numbers may be ignored and the nearest whole number taken. Thus, diborane contains atoms in the ratio 1 B:3 H. The compound's empirical formula is  $\text{BH}_3$ .

Sometimes mass composition is known instead of percentage composition. To determine the empirical formula in this case, convert mass composition to composition in moles. Then calculate the smallest whole-number mole ratio of atoms. This process is shown in Sample Problem 7-13.

### SAMPLE PROBLEM 7-12

**Quantitative analysis shows that a compound contains 32.38% sodium, 22.65% sulfur, and 44.99% oxygen. Find the empirical formula of this compound.**

#### SOLUTION

- 1 **ANALYZE** **Given:** percentage composition: 32.38% Na, 22.65% S, and 44.99% O  
**Unknown:** empirical formula
- 2 **PLAN** percentage composition  $\longrightarrow$  mass composition  $\longrightarrow$  composition in moles  
 $\longrightarrow$  smallest whole-number mole ratio of atoms
- 3 **COMPUTE** Mass composition (mass of each element in 100.0 g sample): 32.38 g Na, 22.65 g S, 44.99 g O  
  
Composition in moles:  $32.38 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 1.408 \text{ mol Na}$   
 $22.65 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.7063 \text{ mol S}$   
 $44.99 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.812 \text{ mol O}$   
  
Smallest whole-number mole ratio of atoms:  
The compound contains atoms in the ratio 1.408 mol Na:0.7063 mol S:2.812 mol O. To find the smallest whole-number mole ratio, divide each value by the smallest number in the ratio.  
$$\frac{1.408 \text{ mol Na}}{0.7063} : \frac{0.7063 \text{ mol S}}{0.7063} : \frac{2.812 \text{ mol O}}{0.7063} = 1.993 \text{ mol Na} : 1 \text{ mol S} : 3.981 \text{ mol O}$$
  
  
Rounding each number in the ratio to the nearest whole number yields a mole ratio of 2 mol Na:1 mol S:4 mol O. The empirical formula of the compound is  $\text{Na}_2\text{SO}_4$ .
- 4 **EVALUATE** Calculating the percentage composition of the compound based on the empirical formula determined in the problem reveals a percentage composition of 32.37% Na, 22.58% S, and 45.05% O. These values agree reasonably well with the given percentage composition.

## SAMPLE PROBLEM 7-13

Analysis of a 10.150 g sample of a compound known to contain only phosphorus and oxygen indicates a phosphorus content of 4.433 g. What is the empirical formula of this compound?

### SOLUTION

#### 1 ANALYZE

**Given:** sample mass = 10.150 g  
phosphorus mass = 4.433 g  
**Unknown:** empirical formula

#### 2 PLAN

Mass composition  $\longrightarrow$  composition in moles  $\longrightarrow$  smallest whole-number ratio of atoms

#### 3 COMPUTE

The mass of oxygen is found by subtracting the phosphorus mass from the sample mass.

$$\text{sample mass} - \text{phosphorus mass} = 10.150 \text{ g} - 4.433 \text{ g} = 5.717 \text{ g}$$

Mass composition: 4.433 g P, 5.717 g O

Composition in moles:

$$4.433 \text{ g } \cancel{\text{P}} \times \frac{1 \text{ mol P}}{30.97 \text{ g } \cancel{\text{P}}} = 0.1431 \text{ mol P}$$

$$5.717 \text{ g } \cancel{\text{O}} \times \frac{1 \text{ mol O}}{16.00 \text{ g } \cancel{\text{O}}} = 0.3573 \text{ mol O}$$

Smallest whole-number mole ratio of atoms:

$$\frac{0.1431 \text{ mol P}}{0.1431} : \frac{0.3573 \text{ mol O}}{0.1431}$$

$$1 \text{ mol P} : 2.497 \text{ mol O}$$

The number of O atoms is not close to a whole number. But if we multiply each number in the ratio by 2, then the number of O atoms becomes 4.994 mol, which is close to 5 mol. The simplest whole-number mole ratio of P atoms to O atoms is 2:5. The compound's empirical formula is  $\text{P}_2\text{O}_5$ .

#### 4 EVALUATE

The arithmetic is correct, significant figures have been used correctly, and units cancel as desired. The formula is reasonable because +5 is a common oxidation state of phosphorus.

### PRACTICE

1. A compound is found to contain 63.52% iron and 36.48% sulfur. Find its empirical formula.

*Answer*  
 $\text{FeS}$

2. Find the empirical formula of a compound found to contain 26.56% potassium, 35.41% chromium, and the remainder oxygen.

*Answer*  
 $\text{K}_2\text{Cr}_2\text{O}_7$

3. Analysis of 20.0 g of a compound containing only calcium and bromine indicates that 4.00 g of calcium are present. What is the empirical formula of the compound formed?

*Answer*  
 $\text{CaBr}_2$

## Calculation of Molecular Formulas

Remember that the *empirical formula* contains the smallest possible whole numbers that describe the atomic ratio. The *molecular formula* is the actual formula of a molecular compound. An empirical formula may or may not be a correct molecular formula. For example, diborane's empirical formula is  $\text{BH}_3$ . Any multiple of  $\text{BH}_3$ , such as  $\text{B}_2\text{H}_6$ ,  $\text{B}_3\text{H}_9$ ,  $\text{B}_4\text{H}_{12}$ , and so on, represents the same ratio of B atoms to H atoms. The molecular compounds ethene,  $\text{C}_2\text{H}_4$ , and cyclopropane,  $\text{C}_3\text{H}_6$ , also share an identical atomic ratio (2 H:1 C), yet they are very different substances. How is the correct formula of a molecular compound found from an empirical formula?

The relationship between a compound's empirical formula and its molecular formula can be written as follows.

$$x(\text{empirical formula}) = \text{molecular formula}$$

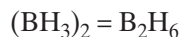
The number represented by  $x$  is a whole-number multiple indicating the factor by which the subscripts in the empirical formula must be multiplied to obtain the molecular formula. (The value of  $x$  is sometimes 1.) The formula masses have a similar relationship.

$$x(\text{empirical formula mass}) = \text{molecular formula mass}$$

To determine the molecular formula of a compound, you must know the compound's formula mass. For example, experimentation shows the formula mass of diborane to be 27.67 amu. The formula mass for the empirical formula,  $\text{BH}_3$ , is 13.84 amu. Dividing the experimental formula mass by the empirical formula mass gives the value of  $x$  for diborane.

$$x = \frac{27.67 \text{ amu}}{13.84 \text{ amu}} = 2.000$$

The molecular formula of diborane is therefore  $\text{B}_2\text{H}_6$ .



Recall that a compound's molecular formula mass is numerically equal to its molar mass, so a compound's molecular formula can also be found given the compound's empirical formula and its molar mass.

### SAMPLE PROBLEM 7-14

In Sample Problem 7-13, the empirical formula of a compound of phosphorus and oxygen was found to be  $\text{P}_2\text{O}_5$ . Experimentation shows that the molar mass of this compound is 283.89 g/mol. What is the compound's molecular formula?

#### SOLUTION

1

#### ANALYZE

Given: empirical formula

Unknown: molecular formula

**2 PLAN**

$$x(\text{empirical formula}) = \text{molecular formula}$$

$$x = \frac{\text{molecular formula mass}}{\text{empirical formula mass}}$$

**3 COMPUTE**

Molecular formula mass is numerically equal to molar mass. Thus, changing the g/mol unit of the compound's molar mass to amu yields the compound's molecular formula mass.

$$\text{molecular molar mass} = 283.89 \text{ g/mol}$$

$$\text{molecular formula mass} = 283.89 \text{ amu}$$

The empirical formula mass is found by adding the masses of each of the atoms indicated in the empirical formula.

$$\text{mass of phosphorus atom} = 30.97 \text{ amu}$$

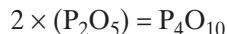
$$\text{mass of oxygen atom} = 16.00 \text{ amu}$$

$$\text{empirical formula mass of } \text{P}_2\text{O}_5 = 2 \times 30.97 \text{ amu} + 5 \times 16.00 \text{ amu} = 141.94 \text{ amu}$$

Dividing the experimental formula mass by the empirical formula mass gives the value of  $x$ . The formula mass is numerically equal to the molar mass.

$$x = \frac{283.89 \text{ amu}}{141.94 \text{ amu}} = 2.0001$$

The compound's molecular formula is therefore  $\text{P}_4\text{O}_{10}$ .

**4 EVALUATE**

Checking the arithmetic shows that it is correct.

**PRACTICE**

1. Determine the molecular formula of the compound with an empirical formula of CH and a formula mass of 78.110 amu. *Answer*  
 $\text{C}_6\text{H}_6$
2. A sample of a compound with a formula mass of 34.00 amu is found to consist of 0.44 g H and 6.92 g O. Find its molecular formula. *Answer*  
 $\text{H}_2\text{O}_2$

**SECTION REVIEW**

1. A compound is found to contain 36.48% Na, 25.41% S, and 38.11% O. Find its empirical formula.
2. Find the empirical formula of a compound that contains 53.70% iron and 46.30% sulfur.
3. Analysis of a compound indicates that it contains 1.04 g K, 0.70 g Cr, and 0.86 g O. Find its empirical formula.
4. If 4.04 g of N combine with 11.46 g O to produce a compound with a formula mass of 108.0 amu, what is the molecular formula of this compound?
5. The molar mass of a compound is 92 g/mol. Analysis of a sample of the compound indicates that it contains 0.606 g N and 1.390 g O. Find its molecular formula.



# CHAPTER 7 REVIEW

## CHAPTER SUMMARY

- 7-1**
- A positive monatomic ion is identified simply by the name of the appropriate element. A negative monatomic ion is named by dropping parts of the ending of the element's name and adding *-ide* to the root.
  - The charge of each ion in an ionic compound may be used to determine the simplest chemical formula for the compound.
  - Compounds composed of two different elements are known as binary compounds.
  - Binary ionic compounds are named by combining the names of the positive and negative ions. Compounds containing polyatomic ions are named in the same manner.
  - The old system of naming binary molecular compounds uses prefixes. The new system, known as the Stock system, uses oxidation numbers.

### Vocabulary

binary compounds (206)  
monatomic ions (204)

nomenclature (206)

oxyanions (209)

salt (215)

- 7-2**
- Oxidation numbers, or oxidation states, are assigned to atoms in compounds according to a set of specific rules. Oxidation numbers are useful in naming compounds, in writing formulas, and in balancing chemical equations.
  - Compounds containing elements that have more than one oxidation state are named using the Stock system of nomenclature.
  - Stock-system names and prefix-system names are used interchangeably for many molecular compounds.
  - In many molecular compounds, oxidation numbers of each element in the compound may be used to determine the compound's simplest chemical formula.
  - By knowing oxidation numbers, we can name compounds without knowing whether they are ionic or molecular.

### Vocabulary

oxidation numbers (216)

oxidation states (216)

- 7-3**
- Formula mass, molar mass, and percentage composition can be calculated from the chemical formula for a compound.
  - The percentage composition of a compound is the percentage by mass of each element in the compound.
  - Molar mass can be used as a conversion factor between an amount in moles and a mass in grams of a given compound or element.

### Vocabulary

formula mass (221)

percentage composition (227)

- 7-4**
- An empirical formula shows the simplest whole-number ratio of atoms in a given compound.
  - Each molecule of a molecular compound contains a whole-number multiple of the atoms in the empirical formula. In some cases this whole-number multiple is equal to 1.
  - Empirical formulas indicate how many atoms of each element are combined in the simplest unit of a chemical compound.
  - A molecular formula can be found from the empirical formula if the molar mass is measured.

### Vocabulary

empirical formula (229)

## REVIEWING CONCEPTS

- What are monatomic ions?
  - Give three examples of monatomic ions. (7-1)
- What is the difference between the nitrite ion and the nitrate ion? (7-1)
- Using only the periodic table, write the symbol of the ion most typically formed by each of the following elements:
  - K
  - Ca
  - S
  - Cl
  - Ba
  - Br
 (7-1)
- Write the formulas and indicate the charges for each of the following ions:
  - sodium ion
  - aluminum ion
  - chloride ion
  - nitride ion
  - iron(II) ion
  - iron(III) ion
 (7-1)
- Name each of the following monatomic ions:
 

a. $K^+$	d. $Cl^-$
b. $Mg^{2+}$	e. $O^{2-}$
c. $Al^{3+}$	f. $Ca^{2+}$

 (7-1)
- Write formulas for the binary ionic compounds formed between the following elements. (Hint: See Sample Problem 7-1.)
  - sodium and iodine
  - calcium and sulfur
  - zinc and chlorine
  - barium and fluorine
  - lithium and oxygen
 (7-1)
- Give the name of each of the following binary ionic compounds. (Hint: See Sample Problem 7-2.)
 

a. KCl	c. $Li_2O$
b. $CaBr_2$	d. $MgCl_2$

 (7-1)
- Write the formulas and give the names of the compounds formed by the following ions:
  - $Cr^{2+}$  and  $F^-$
  - $Ni^{2+}$  and  $O^{2-}$
  - $Fe^{3+}$  and  $O^{2-}$
 (7-1)
- In naming and writing formulas for binary molecular compounds, what determines the order in which the component elements are written? (7-1)
- Name the following binary molecular compounds according to the prefix system. (Hint: See Sample Problem 7-4.)
 

a. $CO_2$	d. $SeF_6$
b. $CCl_4$	e. $As_2O_5$
c. $PCl_5$	

 (7-1)
- Write formulas for each of the following binary molecular compounds. (Hint: See Sample Problem 7-4.)
  - carbon tetrabromide
  - silicon dioxide
  - tetraphosphorus decoxide
  - diarsenic trisulfide
 (7-1)
- Distinguish between binary acids and oxyacids, and give two examples of each. (7-1)
- What is a salt?
  - Give two examples of salts.
 (7-1)
- Name each of the following acids:
 

a. HF	d. $H_2SO_4$
b. HBr	e. $H_3PO_4$
c. $HNO_3$	

 (7-1)
- Give the molecular formula of each of the following acids:
  - sulfurous acid
  - chloric acid
  - hydrochloric acid
  - hypochlorous acid
  - perchloric acid
  - carbonic acid
  - acetic acid
 (7-3)
- Name each of the following ions according to the Stock system:
 

a. $Fe^{2+}$	d. $Pb^{4+}$
b. $Fe^{3+}$	e. $Sn^{2+}$
c. $Pb^{2+}$	f. $Sn^{4+}$

 (7-2)
- Name each of the binary molecular compounds in item 11 using the Stock system. (7-2)
- Write formulas for each of the following compounds:
  - phosphorus(III) iodide
  - sulfur(II) chloride
  - carbon(IV) sulfide
  - nitrogen(V) oxide
 (7-2)

19. a. What are oxidation numbers?  
b. What useful functions do oxidation numbers serve? (7-2)
20. a. Define *formula mass*.  
b. In what unit is formula mass expressed? (7-3)
21. What is meant by the molar mass of a compound? (7-3)
22. What three types of information are needed in order to find an empirical formula from percentage composition data? (7-4)
23. What is the relationship between the empirical formula and the molecular formula of a compound? (7-4)

## PROBLEMS

### Nomenclature and Chemical Formulas

24. Write the formula and charge for each of the following ions:
  - a. ammonium ion
  - b. acetate ion
  - c. hydroxide ion
  - d. carbonate ion
  - e. sulfate ion
  - f. phosphate ion
  - g. copper(II) ion
  - h. tin(II) ion
  - i. iron(III) ion
  - j. copper(I) ion
  - k. mercury(I) ion
  - l. mercury(II) ion
25. Name each of the following ions:
 

a. $\text{NH}_4^+$	f. $\text{CO}_3^{2-}$
b. $\text{ClO}_3^-$	g. $\text{PO}_4^{3-}$
c. $\text{OH}^-$	h. $\text{CH}_3\text{COO}^-$
d. $\text{SO}_4^{2-}$	i. $\text{HCO}_3^-$
e. $\text{NO}_3^-$	j. $\text{CrO}_4^{2-}$
26. Write formulas for each of the following compounds:
  - a. sodium fluoride
  - b. calcium oxide
  - c. potassium sulfide
  - d. magnesium chloride
  - e. aluminum bromide
  - f. lithium nitride
  - g. iron(II) oxide

### Oxidation Number and Stock System

27. Name each of the following ionic compounds using the Stock system:

- a. NaCl
- b. KF
- c. CaS
- d.  $\text{Co}(\text{NO}_3)_2$
- e.  $\text{FePO}_4$
- f.  $\text{Hg}_2\text{SO}_4$
- g.  $\text{Hg}_3(\text{PO}_4)_2$

28. Assign oxidation numbers to each atom in the following compounds. (Hint: See Sample Problem 7-5.)

- a. HI
- b.  $\text{PBr}_3$
- c.  $\text{GeS}_2$
- d. KH
- e.  $\text{As}_2\text{O}_5$
- f.  $\text{H}_3\text{PO}_4$

29. Assign oxidation numbers to each atom in the following ions. (Hint: See Sample Problem 7-5.)

- a.  $\text{NO}_3^-$
- b.  $\text{ClO}_4^-$
- c.  $\text{PO}_4^{3-}$
- d.  $\text{Cr}_2\text{O}_7^{2-}$
- e.  $\text{CO}_3^{2-}$

### Mole Relationships and Percentage Composition

30. Determine the formula mass of each of the following compounds or ions. (Hint: See Sample Problem 7-6.)

- a. glucose,  $\text{C}_6\text{H}_{12}\text{O}_6$
- b. calcium acetate,  $\text{Ca}(\text{CH}_3\text{COO})_2$
- c. the ammonium ion,  $\text{NH}_4^+$
- d. the chlorate ion,  $\text{ClO}_3^-$

31. Determine the number of moles of each type of monatomic or polyatomic ion in one mole of the following compounds. For each polyatomic ion, determine the number of moles of each atom present in one mole of the ion.

- a.  $\text{KNO}_3$
- b.  $\text{Na}_2\text{SO}_4$
- c.  $\text{Ca}(\text{OH})_2$
- d.  $(\text{NH}_4)_2\text{SO}_3$
- e.  $\text{Ca}_3(\text{PO}_4)_2$
- f.  $\text{Al}_2(\text{CrO}_4)_3$

32. Determine the molar mass of each compound listed in item 31. (Hint: See Sample Problem 7-7.)
33. Determine the number of moles of compound in each of the following samples. (Hint: See Sample Problem 7-9.)
  - a. 4.50 g  $\text{H}_2\text{O}$
  - b. 471.6 g  $\text{Ba}(\text{OH})_2$
  - c. 129.68 g  $\text{Fe}_3(\text{PO}_4)_2$
34. Determine the percentage composition of each of the following compounds. (Hint: See Sample Problem 7-10.)
  - a.  $\text{NaCl}$
  - b.  $\text{AgNO}_3$
  - c.  $\text{Mg}(\text{OH})_2$
35. Determine the percentage by mass of water in the hydrate  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ . (Hint: See Sample Problem 7-11.)
36. Determine the empirical formula of a compound containing 63.50% silver, 8.25% nitrogen, and the remainder oxygen. (Hint: See Sample Problem 7-12.)
37. Determine the empirical formula of a compound found to contain 52.11% carbon, 13.14% hydrogen, and 34.75% oxygen.
38. What is the molecular formula of the molecule that has an empirical formula of  $\text{CH}_2\text{O}$  and a molar mass of 120.12 g/mol?
39. A compound with a formula mass of 42.08 amu is found to be 85.64% carbon and 14.36% hydrogen by mass. Find its molecular formula.
42. What is the mass in grams of each of the following samples?
  - a. 1.000 mol  $\text{NaCl}$
  - b. 2.000 mol  $\text{H}_2\text{O}$
  - c. 3.500 mol  $\text{Ca}(\text{OH})_2$
  - d. 0.625 mol  $\text{Ba}(\text{NO}_3)_2$
43. Determine the formula mass and molar mass of each of the following compounds:
  - a.  $\text{XeF}_4$
  - b.  $\text{C}_{12}\text{H}_{24}\text{O}_6$
  - c.  $\text{Hg}_2\text{I}_2$
  - d.  $\text{CuCN}$
44. Write the chemical formulas for the following compounds:
  - a. aluminum fluoride
  - b. magnesium oxide
  - c. vanadium(V) oxide
  - d. cobalt(II) sulfide
  - e. strontium bromide
  - f. sulfur trioxide
45. How many atoms of each element are contained in a single formula unit of iron(III) formate,  $\text{Fe}(\text{CHO}_2)_3 \cdot \text{H}_2\text{O}$ ? What percentage by mass of the compound is water?
46. Name each of the following acids, and assign oxidation numbers to the atoms in each:
 

a. $\text{HNO}_2$	c. $\text{H}_2\text{CO}_3$
b. $\text{H}_2\text{SO}_3$	d. $\text{HI}$
47. Determine the percentage composition of the following compounds:
  - a.  $\text{NaClO}$
  - b.  $\text{H}_2\text{SO}_3$
  - c.  $\text{C}_2\text{H}_5\text{COOH}$
  - d.  $\text{BeCl}_2$
48. Name each of the following binary compounds:
 

a. $\text{MgI}_2$	e. $\text{SO}_2$
b. $\text{NaF}$	f. $\text{PBr}_3$
c. $\text{CS}_2$	g. $\text{CaCl}_2$
d. $\text{N}_2\text{O}_4$	h. $\text{AgI}$
49. Assign oxidation numbers to each atom in the following molecules and ions:
 

a. $\text{CO}_2$	e. $\text{H}_2\text{O}_2$
b. $\text{NH}_4^+$	f. $\text{P}_4\text{O}_{10}$
c. $\text{MnO}_4^-$	g. $\text{OF}_2$
d. $\text{S}_2\text{O}_3^{2-}$	

### MIXED REVIEW

40. Chemical analysis of citric acid shows that it contains 37.51% C, 4.20% H, and 58.29% O. What is its empirical formula?
41. Name each of the following compounds using the Stock system:
 

a. $\text{LiBr}$	f. $\text{Fe}_2\text{O}_3$
b. $\text{Sn}(\text{NO}_3)_2$	g. $\text{AgNO}_3$
c. $\text{FeCl}_2$	h. $\text{Fe}(\text{OH})_2$
d. $\text{MgO}$	i. $\text{CrF}_2$
e. $\text{KOH}$	

50. A 175.0 g sample of a compound contains 56.15 g C, 9.43 g H, 74.81 g O, 13.11 g N, and 21.49 g Na. What is its empirical formula?

### CRITICAL THINKING

51. **Analyzing Information** Sulfur trioxide is produced in the atmosphere through a reaction of sulfur dioxide and oxygen. Sulfur dioxide is a primary air pollutant. List all of the chemical information you can by analyzing the formula for sulfur trioxide.
52. **Analyzing Data** In the laboratory, a sample of pure nickel was placed in a clean, dry, weighed crucible. The crucible was heated so that the nickel would react with the oxygen in the air. After the reaction appeared complete, the crucible was allowed to cool and the mass was determined. The crucible was reheated and allowed to cool. Its mass was then determined again to be certain that the reaction was complete. The following data were collected:

Mass of crucible	= 30.02 g
Mass of nickel and crucible	= 31.07 g
Mass of nickel oxide and crucible	= 31.36 g

Determine the following information based on the data given above:

Mass of nickel	=
Mass of nickel oxide	=
Mass of oxygen	=

Based on your calculations, what is the empirical formula for the nickel oxide?

mass of each element in the formula. It then can be used to find the molar masses of various compounds.

Go to Appendix C. If you are using a TI 83 Plus, you can download the program and data and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. Remember that you will need to name the program and check the display, as explained in Appendix C. You will then be ready to run the program. After you have graphed the data, answer these questions.

Note: All answers are written with 2 significant digits beyond the decimal point.

- What is the molar mass of  $\text{BaTiO}_3$ ?
- What is the molar mass of  $\text{PbCl}_2$ ?
- What is the molar mass of  $\text{NH}_4\text{NO}_3$ ?



### HANDBOOK SEARCH

54. Review the common reactions of Group 1 metals in the *Elements Handbook* and answer the following:
- Some of the Group 1 metals react with oxygen to form superoxides. Write the formulas for these compounds.
  - What is the charge on each cation for the formulas you wrote in (a)?
  - How does the charge on the anion vary for oxides, peroxides, and superoxides?
55. Review the common reactions of Group 2 metals in the *Elements Handbook* and answer the following:
- Some of the Group 2 metals react with oxygen to form oxides. Write the formulas for these compounds.
  - Some of the Group 2 metals react with oxygen to form peroxides. Write the formulas for these compounds.
  - Some of the Group 2 metals react with nitrogen to form nitrides. Write the formulas for these compounds.



### TECHNOLOGY & LEARNING

53. **Graphing Calculator** Calculate the Molar Mass of a Compound

The graphing calculator can run a program that calculates the molar mass of a compound given the chemical formula for the compound. This program will prompt for the number of elements in the formula, the number of atoms of each element in the formula, and the atomic



- d. Most Group 2 elements form hydrides.  
What is hydrogen's oxidation state in these compounds?

56. Review the analytical tests for transition metals in the *Elements Handbook* and answer the following:
- Determine the oxidation state of each metal in the precipitates shown for cadmium, zinc, and lead.
  - Determine the oxidation state of each metal in the complex ions shown for iron, manganese, and cobalt.
  - The copper compound shown is called a coordination compound. The ammonia shown in the formula exists as molecules with no charge. Determine copper's oxidation state in this compound.
57. Review the common reactions of Group 15 elements in the *Elements Handbook* and answer the following:
- Write formulas for each of the oxides listed for the Group 15 elements.
  - Determine nitrogen's oxidation state in the oxides listed in (a).

## RESEARCH & WRITING

58. **Nomenclature** Biologists who name newly discovered organisms use a system that is structured very much like the one used by chemists in naming compounds. The system used by biologists is called the Linnaeus system, after its creator, Carolus Linnaeus. Research this system

in a biology textbook, and then note similarities and differences between the Linnaeus system and chemical nomenclature.

59. **Common Chemicals** Find out the systematic chemical name and write the chemical formula for each of the following common compounds:
- |                     |                 |
|---------------------|-----------------|
| a. baking soda      | d. limestone    |
| b. milk of magnesia | e. lye          |
| c. Epsom salts      | f. wood alcohol |

## ALTERNATIVE ASSESSMENT

60. **Performance Assessment** Your teacher will supply you with a note card with one of the following formulas on it:  $\text{NaCH}_3\text{COO} \cdot 3\text{H}_2\text{O}$ ,  $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ ,  $\text{LiC}_2\text{H}_3\text{O}_2 \cdot 2\text{H}_2\text{O}$ , or  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ . Design an experiment to determine the percentage of water by mass in the hydrated salt assigned to you. Be sure to explain what steps you will take to ensure that the salt is completely dry. If your teacher approves your design, obtain the salt and perform the experiment. What percentage of water does the salt contain?
61. Both ammonia,  $\text{NH}_3$ , and ammonium nitrate,  $\text{NH}_4\text{NO}_3$ , are used in fertilizers as a source of nitrogen. Which compound has the higher percentage of nitrogen? Research the physical properties of both compounds, and find out how each is manufactured and used. Explain why each compound has its own particular application. (Consider factors such as the cost of raw ingredients, the ease of manufacture, shipping costs, and so forth.)