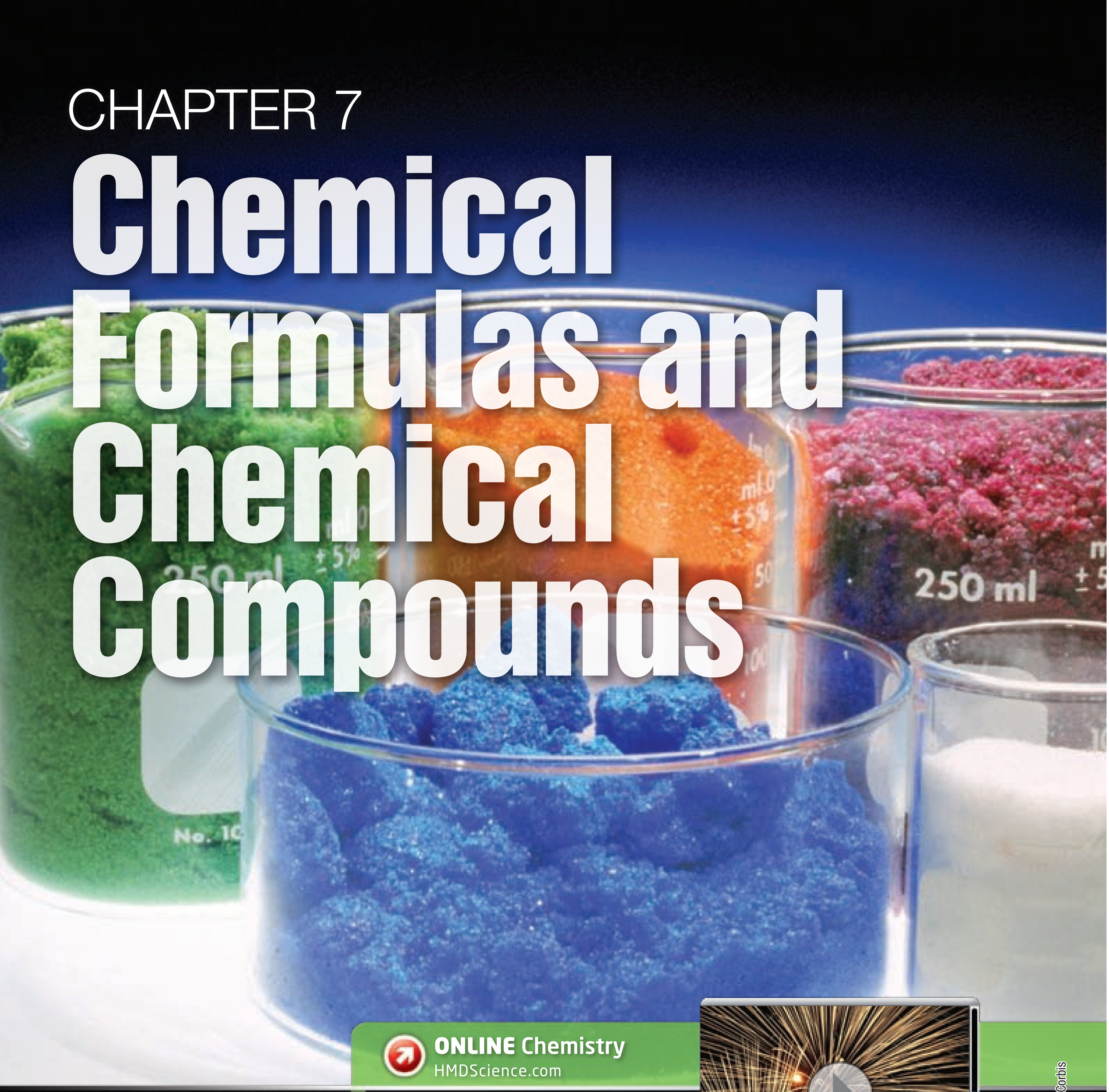
CHAPTER 7

Chemical Formulas and Chemical Compounds



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Formulas and Compounds

Chemical Names and Formulas

SECTION 1

Main Ideas

- Formulas tell the number and kinds of atoms in a compound.
- Monatomic ions are made of only one type of atom.
- Binary compounds contain atoms of two elements.
- Some covalent compounds are a network with no single molecules.
- Acids are solutions of water and a special type of compound.

VIRGINIA STANDARDS

CH.3 The student will investigate and understand how conservation of energy and matter is expressed in chemical formulas and balanced equations. Key concepts include:
CH.3.a nomenclature.
CH.3.c writing chemical formulas.
CH.3.EKS-1; CH.3.EKS-5; CH.3.EKS-6

Key Terms

monatomic ion
binary compound
nomenclature
oxyanion
salt

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 writing chemical formulas. In this chapter, you will be introduced to some of the rules used to identify simple chemical compounds.

MAIN IDEA

Formulas tell the number and kinds of atoms in a compound.

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.)

$$C_8H_{18}$$

Subscript indicates that there are 8 carbon atoms in a molecule of octane.

Subscript indicates that there are 18 hydrogen atoms in a molecule of octane.

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 the next page.

$Al_2(SO_4)_3$

Subscript 2 refers to 2 aluminum atoms.

Subscript 4 refers to 4 oxygen atoms in sulfate ion.

Subscript 3 refers to everything inside parentheses, giving 3 sulfate ions, with a total of 3 sulfur atoms and 12 oxygen atoms.

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.

MAIN IDEA

Monatomic ions are made of only one type of atom.

CHECK FOR UNDERSTANDING

Describe What characteristic of atoms determines whether they will exchange valence electrons (form an ionic compound) or share them (form a covalent compound)? Describe how this happens. Refer to the chapter "Chemical Bonding" for help.

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 Na^+. Group 2 metals lose two electrons to give 2+ cations, such as 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, 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+ and 2+ cations. Iron and chromium each form 2+ cations as well as 3+ cations. And vanadium forms 2+, 3+, and 4+ cations.

Examples of Cations

K^+
Potassium cation

Mg^{2+}
Magnesium cation

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 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 on the next page.

The names and symbols of the common monatomic cations and anions are organized according to their charges in **Figure 1.1.** 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.

Examples of Anions

Element	Anion
F Fluor*ine*	F^- Fluor*ide* anion
N Nitro*gen*	N^{3-} Nitr*ide* anion

FIGURE 1.1

SOME COMMON MONATOMIC IONS

Main-group elements

1+		2+		3+			
lithium	Li^+	beryllium	Be^{2+}	aluminum	Al^{3+}		
sodium	Na^+	magnesium	Mg^{2+}				
potassium	K^+	calcium	Ca^{2+}				
rubidium	Rb^+	strontium	Sr^{2+}				
cesium	Cs^+	barium	Ba^{2+}				
1–		**2–**		**3–**			
fluoride	F^-	oxide	O^{2-}	nitride	N^{3-}		
chloride	Cl^-	sulfide	S^{2-}	phosphide	P^{3-}		
bromide	Br^-						
iodide	I^-						

d-Block elements and others with multiple ions

1+		2+		3+		4+	
copper(I)	Cu^+	vanadium(II)	V^{2+}	vanadium(III)	V^{3+}	vanadium(IV)	V^{4+}
silver	Ag^+	chromium(II)	Cr^{2+}	chromium(III)	Cr^{3+}	tin(IV)	Sn^{4+}
		manganese(II)	Mn^{2+}	iron(III)	Fe^{3+}	lead(IV)	Pb^{4+}
		iron(II)	Fe^{2+}	cobalt(III)	Co^{3+}		
		cobalt(II)	Co^{2+}				
		nickel(II)	Ni^{2+}				
		copper(II)	Cu^{2+}				
		zinc	Zn^{2+}				
		cadmium	Cd^{2+}				
		tin(II)	Sn^{2+}				
		mercury(II)	Hg^{2+}				
		lead(II)	Pb^{2+}				

CAREERS IN CHEMISTRY

Pharmacist S.T.E.M.

Pharmacists work in many areas of healthcare. Traditionally, they have dispensed medications in pharmacies. Today, pharmacists work with other health professionals to develop and implement customized drug treatment plans. They advise medical professionals on the side effects of and interactions between medicines. Increasingly, pharmacists are working in nontraditional settings, such as in managed-care companies, pharmaceutical companies, and governmental agencies. In these settings, pharmacists establish drug guidelines, conduct drug reviews, and assist in the research and development of new drugs.

MAIN IDEA

Binary compounds contain atoms of two elements.

Compounds composed of two 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 Mg^{2+} cation. Note that the $^{2+}$ in Mg^{2+} is written as a superscript. Bromine, a halogen, forms the Br^- anion when combined with a metal. In each formula unit of magnesium bromide, two Br^- anions are required to balance the 2+ charge of the Mg^{2+} cation. The compound's formula must therefore indicate one Mg^{2+} cation and two Br^- anions. The symbol for the cation is written first.

Ions combined: Mg^{2+}, Br^-, Br^- *Chemical formula:* $MgBr_2$

Note that the $_2$ in 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 for 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, Al^{3+}, and the oxide ion, O^{2-}, is determined as follows.

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

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

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

$$Al_2^{3+} \quad O_3^{2-}$$

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

$$Al_2O_3$$

CHECK FOR UNDERSTANDING

Describe In your own words, describe how to determine the subscripts of a binary ionic compound by using the oxidation numbers of the ions.

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 for aluminum chloride. Notice that the known relative charges of each ion make specifying ratios in the name unnecessary.

Al_2O_3

Name of cation — Name of anion

aluminum **oxide**

Writing Formulas For Ionic Compounds

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

a. zinc and iodine **b. zinc and sulfur**

SOLVE

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

a. Zn^{2+} I^{-}

b. Zn^{2+} S^{2-}

Cross over the charges to give subscripts.

a. Zn_1^{2+} I_2^{-}

b. Zn_2^{2+} 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 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 ZnS.

Practice

Answers in Appendix E

1. Write formulas for the binary ionic compounds formed between the following elements:
   - **a.** potassium and iodine
   - **b.** magnesium and chlorine
   - **c.** sodium and sulfur
   - **d.** aluminum and sulfur
   - **e.** aluminum and nitrogen
2. Name the binary ionic compounds indicated by the following formulas:
   - **a.** AgCl
   - **b.** ZnO
   - **c.** $CaBr_2$
   - **d.** SrF_2
   - **e.** BaO
   - **f.** $CaCl_2$

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, scientists use the Stock system of nomenclature. It is useful for distinguishing two different compounds formed by the same elements, as the lead oxides in **Figure 1.2.** This system uses a Roman numeral to indicate an ion's charge. The numeral in parentheses is placed *immediately* after the metal name.

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

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

Na^{+}	Ba^{2+}	Al^{3+}
sodium	barium	aluminum

No element forms more than one monatomic anion.

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

$CuCl_2$

Name of cation + Roman numeral indicating charge — Name of anion

copper(II) **chloride**

FIGURE 1.2

Different Cations of a Metal

Different cations of the same metal form different compounds even when they combine with the same anion.

(a) lead(IV) oxide, PbO_2 **(b)** lead(II) oxide, PbO

Naming Ionic Compounds

Sample Problem B **Write the formula and give the name for the compound formed by the ions Cr^{3+} and F^{-}.**

SOLVE

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

$$Cr^{3+} \;\; F^{-}$$

Cross over the charges to give subscripts.

$$Cr_1^{3+} \;\; F_3^{-}$$

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 CrF_3. As **Figure 1.1** 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 Answers in Appendix E

1. Write the formula and give the name for the compounds formed between the following ions:
   - **a.** Cu^{2+} and Br^{-}
   - **b.** Fe^{2+} and O^{2-}
   - **c.** Pb^{2+} and Cl^{-}
   - **d.** Hg^{2+} and S^{2-}
   - **e.** Sn^{2+} and F^{-}
   - **f.** Fe^{3+} and O^{2-}
2. Give the names for the following compounds:
   - **a.** CuO
   - **b.** CoF_3
   - **c.** SnI_4
   - **d.** FeS

Compounds Containing Polyatomic Ions

Figure 1.3 on the next page lists some common polyatomic ions. Most are negatively charged and most are **oxyanions**—**polyatomic ions that contain oxygen.** Some elements can combine with oxygen to form more than one type of oxyanion. For example, nitrogen can form NO_3^- or NO_2^-. The name given a compound containing such an oxyanion depends on the number of oxygen atoms in the oxyanion. The name of the ion with the greater number of oxygen atoms ends in *-ate*. The name of the ion with the smaller number of oxygen atoms ends in *-ite*.

NO_2^-	NO_3^-
nitrite	nitrate

FIGURE 1.3

SOME POLYATOMIC IONS

1+		2+			
ammonium	NH_4^+	dimercury*	Hg_2^{2+}		
1–		**2–**		**3–**	
acetate	CH_3COO^-	carbonate	CO_3^{2-}	arsenate	AsO_4^{3-}
bromate	BrO_3^-	chromate	CrO_4^{2-}	phosphate	PO_4^{3-}
chlorate	ClO_3^-	dichromate	$Cr_2O_7^{2-}$		
chlorite	ClO_2^-	hydrogen phosphate	HPO_4^{2-}		
cyanide	CN^-	oxalate	$C_2O_4^{2-}$		
dihydrogen phosphate	$H_2PO_4^-$	peroxide	O_2^{2-}		
hydrogen carbonate (bicarbonate)	HCO_3^-	sulfate	SO_4^{2-}		
hydrogen sulfate	HSO_4^-	sulfite	SO_3^{2-}		
hydroxide	OH^-				
hypochlorite	ClO^-				
nitrate	NO_3^-				
nitrite	NO_2^-				
perchlorate	ClO_4^-				
permanganate	MnO_4^-				

*The mercury(I) cation exists as two Hg^+ ions joined together by a covalent bond and is written as Hg_2^{2+}.

Sometimes, an element can form more than two types of oxyanions. In this case, the prefix *hypo-* is given to an anion that has one fewer oxygen atom than the -ite anion. The prefix *per-* is given to an anion that has one more oxygen atom than the -ate anion. This nomenclature is illustrated by the four oxyanions formed by chlorine.

ClO^-	ClO_2^-	ClO_3^-	ClO_4^-
hypochlorite	chlorite	chlorate	perchlorate

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*, $AgNO_3$, and *silver nitrite*, $AgNO_2$, respectively. When multiples of a polyatomic ion are present in a compound, the formula for the polyatomic ion is enclosed in parentheses.

Writing Formulas for Ionic Compounds

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Sample Problem C **Write the formula for tin(IV) sulfate.**

SOLVE

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

$$Sn^{4+} \; SO_4^{2-}$$

Cross over the charges to give subscripts. Add parentheses around the polyatomic ion if necessary.

$$Sn_2^{4+} \; (SO_4)_4^{2-}$$

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 $Sn(SO_4)_2$.

Practice

1. Write formulas for the following ionic compounds:
   - **a.** sodium iodide
   - **b.** calcium chloride
   - **c.** potassium sulfide
   - **d.** lithium nitrate
   - **e.** copper(II) sulfate
   - **f.** sodium carbonate
   - **g.** calcium nitrite
   - **h.** potassium perchlorate
2. Give the names for the following compounds:
   - **a.** Ag_2O
   - **b.** $Ca(OH)_2$
   - **c.** $KClO_3$
   - **d.** NH_4OH
   - **e.** $Fe_2(CrO_4)_3$
   - **f.** $KClO$

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 2.

The old system of naming molecular compounds is based on the use of prefixes. For example, the molecular compound CCl_4 is named carbon *tetra*chloride. The prefix *tetra-* indicates that four chloride atoms are present in a single molecule of the compound. The two oxides of carbon, CO and CO_2, are named carbon *mon*oxide and carbon *di*oxide, respectively. These prefix-based names are often the most widely-recognized names for some molecular compounds. However, either naming system is acceptable, unless specified otherwise.

FIGURE 1.4

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-* (from *mono-*) indicates one oxygen atom, and the prefix *di-* indicates two oxygen atoms. The prefixes used to specify the number of atoms in a molecule are listed in **Figure 1.4.**

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

1. The element that has the smaller group number is usually given first. If both elements are in the same group, the element whose period number is greater is given first. The element 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 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. For example, one would write *monoxide* and *pentoxide* instead of *mono-oxide* and *penta-oxide*.

The prefix system is illustrated below.

P_4O_{10}

Prefix needed if first element contributes more than one atom + Name of first element | Prefix indicating number of atoms contributed by second element + Root name of second element + *ide*

tetraphosphorus **decoxide**

In general, the order of nonmetals in binary compound names and formulas is C, P, N, H, S, I, Br, Cl, O, and F.

FIGURE 1.5

BINARY COMPOUNDS OF NITROGEN AND OXYGEN

Formula	Prefix-system name
N_2O	dinitrogen monoxide
NO	nitrogen monoxide
NO_2	nitrogen dioxide
N_2O_3	dinitrogen trioxide
N_2O_4	dinitrogen tetroxide
N_2O_5	dinitrogen pentoxide

The prefix system is illustrated further in **Figure 1.5**, 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_2. No prefix is needed with *nitrogen* because only one atom of nitrogen is present in a molecule of NO_2. On the other hand, according to rule 2, the prefix *di-* in *dioxide* is needed to indicate the presence of two atoms of oxygen. Take a moment to review the prefixes in the other names in **Figure 1.5.**

CHECK FOR UNDERSTANDING

Explain You have learned several ways to name compounds. Explain why it is necessary to have so many different methods of nomenclature.

Naming Binary Molecular Compounds

Sample Problem D **a. Give the name for As_2O_5.**
b. Write the formula for oxygen difluoride.

SOLVE

a. 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*.

b. 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_2.

Practice

Answers in Appendix E

1. Name the following binary molecular compounds:
   - **a.** SO_3
   - **b.** ICl_3
   - **c.** PBr_5
2. Write formulas for the following compounds:
   - **a.** carbon tetraiodide
   - **b.** phosphorus trichloride
   - **c.** dinitrogen trioxide

MAIN IDEA

Some covalent compounds are a network with no single molecules.

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.

SiC	SiO_2	Si_3N_4
silicon carbide	silicon dioxide	trisilicon tetranitride

MAIN IDEA

Acids are solutions of water and a special type of compound.

An *acid* is a distinct type of molecular compound. 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 **Figure 1.6. Figure 1.7** shows some common laboratory acids.

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	H_2SO_4	sulfate	SO_4^{2-}
nitric acid	HNO_3	nitrate	NO_3^-
phosphoric acid	H_3PO_4	phosphate	PO_4^{3-}

FIGURE 1.6

COMMON BINARY ACIDS AND OXYACIDS

HF	hydrofluoric acid	HNO_2	nitrous acid	HClO	hypochlorous acid
HCl	hydrochloric acid	HNO_3	nitric acid	$HClO_2$	chlorous acid
HBr	hydrobromic acid	H_2SO_3	sulfurous acid	$HClO_3$	chloric acid
HI	hydriodic acid	H_2SO_4	sulfuric acid	$HClO_4$	perchloric acid
H_3PO_4	phosphoric acid	CH_3COOH	acetic acid	H_2CO_3	carbonic acid

FIGURE 1.7

Common Acids Some common laboratory acids are nitric acid, hydrochloric acid, and sulfuric acid. Acids should always be handled with care and according to instructions. They can burn holes in skin or clothing.

An ionic compound composed of a cation and the anion from an acid is often referred to as a salt. Table salt, NaCl, contains the anion from hydrochloric acid. Calcium sulfate, $CaSO_4$, 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, H_2CO_3.

HCO_3^-
hydrogen carbonate ion
bicarbonate ion

SECTION 1 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. What is the significance of a chemical formula?

2. Write formulas for the compounds formed between the following:
- **a.** aluminum and bromine
- **b.** sodium and oxygen
- **c.** magnesium and iodine
- **d.** Pb^{2+} and O^{2-}
- **e.** Sn^{2+} and I^-
- **f.** Fe^{3+} and S^{2-}
- **g.** Cu^{2+} and NO_3^-
- **h.** NH_4^+ and SO_4^{2-}

3. Name the following compounds by using the Stock system:
- **a.** NaI
- **b.** MgS
- **c.** CaO
- **d.** K_2S
- **e.** CuBr
- **f.** $FeCl_2$

4. Write formulas for each of the following compounds:
- **a.** sodium hydroxide
- **b.** lead(II) nitrate
- **c.** iron(II) sulfate
- **d.** diphosphorus trioxide
- **e.** carbon diselenide
- **f.** acetic acid
- **g.** chloric acid
- **h.** sulfurous acid

Critical Thinking

5. RELATING IDEAS Draw the Lewis structure, give the name, and predict VSEPR geometry of SCl_2.

SECTION 2

Main Ideas

- Specific rules are used to assign oxidation numbers.
- Many nonmetals have multiple oxidation numbers.

VIRGINIA STANDARDS

CH.1.j The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include: the use of current applications to reinforce chemistry concepts.
CH.2.g The student will investigate and understand that the placement of elements on the periodic table is a function of their atomic structure. The periodic table is a tool used for the investigations of: electron configurations, valence electrons, and oxidation numbers.
CH.3 The student will investigate and understand how conservation of energy and matter is expressed in chemical formulas and balanced equations. Key concepts include:
CH.3.a nomenclature.
CH.3.c writing chemical formulas.
CH.3.EKS-2; CH.3.EKS-5

Oxidation Numbers

Key Terms
oxidation number
oxidation state

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, they are quite arbitrary in some cases. However, oxidation numbers are useful in naming compounds, in writing formulas, and in balancing chemical equations. And, as is discussed in the chapter "Chemical Equations and Reactions", they are helpful in studying certain reactions.

MAIN IDEA

Specific rules are used to assign 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_2, phosphorus, P_4, and sulfur, S_8, 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 forms peroxides, such as H_2O_2, in which its oxidation number is −1, and when it forms compounds with fluorine, such as OF_2, 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.

9. A monatomic ion has an oxidation number equal to the charge of the ion. For example, the ions Na+, Ca^2+, and 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, H_2O. Both compounds have polar-covalent bonds. In HF, the fluorine atom should have an oxidation number of −1 (see Rule 3 on the previous page). Rule 5 tells us that hydrogen should have an oxidation number of +1. This makes sense because fluorine is more electronegative than hydrogen, and in a polar-covalent bond, shared electrons are assumed to belong to the more-electronegative element. For water, Rules 4 and 5 tell us that the oxidation number of oxygen should be −2, and the oxidation number of each hydrogen atom should be +1. Again, oxygen is more electronegative than hydrogen, so the shared electrons are assumed to belong to oxygen.

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

Oxidation Numbers

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

a. UF_6

b. H_2SO_4

c. ClO_3^-

SOLVE

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.

$$\mathrm{U}\overset{-1}{\mathrm{F}}_6$$

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

$$\mathrm{U}\underset{-6}{\overset{-1}{\mathrm{F}}}_6$$

The compound 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.

$$\underset{+6}{\mathrm{U}}\underset{-6}{\overset{-1}{\mathrm{F}}}_6$$

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.

Continued

$$\overset{+6\,-1}{\underset{+6\,-6}{UF_6}}$$

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_2SO_4 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.

$$\overset{+1}{\underset{+2}{H_2}}S\overset{-2}{\underset{-8}{O_4}}$$

The sum of the oxidation numbers must equal zero, and there is only one sulfur atom in each molecule of H_2SO_4. Because $(+2) + (-8) = -6$, the oxidation number of each sulfur atom must be +6.

c. To assign oxidation numbers to the elements in ClO_3^-, 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.

$$\overset{+5}{\underset{+5}{Cl}}\overset{-2}{\underset{-6}{O_3^-}}$$

Practice

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

a. HCl
b. CF_4
c. PCl_3
d. SO_2
e. HNO_3
f. KH
g. P_4O_{10}
h. $HClO_3$
i. N_2O_5
j. $GeCl_2$

MAIN IDEA

Many nonmetals have multiple oxidation numbers.

As shown in **Figure 2.1,** many nonmetals can have more than one oxidation number. (A more extensive list of oxidation numbers is given in Appendix Table B-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_2 or SO_3. Both are known compounds. Of course, a formula must represent facts. Oxidation numbers alone cannot be used to prove the existence of a compound.

FIGURE 2.1

COMMON OXIDATION NUMBERS OF SOME NONMETALS THAT HAVE VARIABLE OXIDATION STATES*		
Group 14	carbon	−4, +2, +4
Group 15	nitrogen phosphorus	−3, +1, +2, +3, +4, +5 −3, +3, +5
Group 16	sulfur	−2, +4, +6
Group 17	chlorine bromine iodine	−1, +1, +3, +5, +7 −1, +1, +3, +5, +7 −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.		

CHECK FOR UNDERSTANDING
Explain Why do some elements have multiple oxidation states?

In Section 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, SO_2 and 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
PCl_3	phosphorus trichloride	phosphorus(III) chloride
PCl_5	phosphorus pentachloride	phosphorus(V) chloride
N_2O	dinitrogen monoxide	nitrogen(I) oxide
NO	nitrogen monoxide	nitrogen(II) oxide
PbO_2	lead dioxide	lead(IV) oxide
Mo_2O_3	dimolybdenum trioxide	molybdenum(III) oxide

SECTION 2 FORMATIVE ASSESSMENT

Reviewing Main Ideas

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

a. HF
b. CI_4
c. H_2O
d. PI_3
e. CS_2
f. Na_2O_2
g. H_2CO_3
h. NO_2^-
i. SO_4^{2-}

2. Name each of the following binary molecular compounds according to the Stock system:

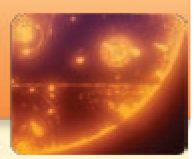
a. CI_4
b. SO_3
c. As_2S_3
d. NCl_3

Critical Thinking

3. DRAWING CONCLUSIONS Determine the oxidation numbers for iron oxide, Fe_3O_4. (Recall that oxidation numbers are integers.)

WHY IT MATTERS





Mass Spectrometry: Identifying Molecules

S.T.E.M.

Tests for locating oil deposits in the ground and detecting dioxins in our food supply are commonly performed today. These tests can be performed by using a technique known as *mass spectrometry*. Mass spectrometry is now used in many fields, such as medicine, chemistry, forensic science, and astronomy.

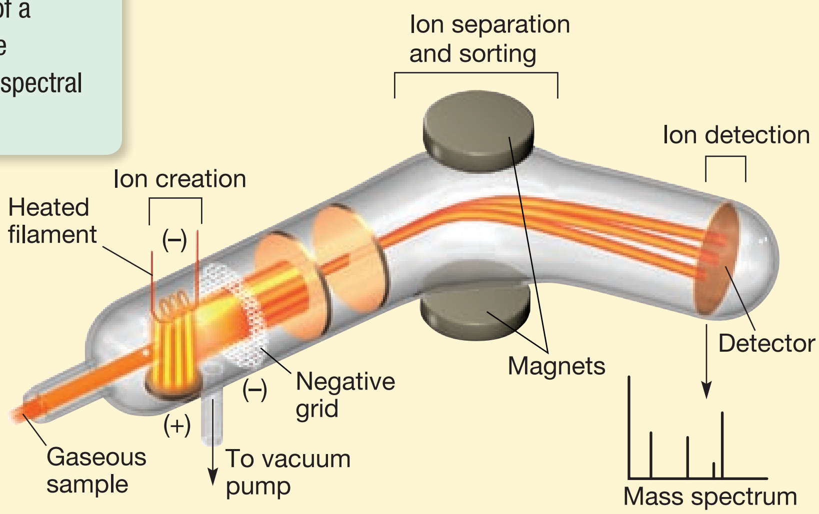
What is mass spectrometry? It is the most accurate technique available to measure the mass of an individual molecule or atom. Knowing the molecular mass is an essential part of identifying an unknown compound and determining the structure of a molecule of the compound. As the diagram of a mass spectrometer shows, the molecules in a gaseous sample are converted into ions. The ions then are separated and sorted according to their mass-to-charge ratio by a combination of electric and magnetic fields. The fields cause the ions' trajectories to change based on the ions' masses and charges. Then, the sorted ions are detected, and a mass spectrum is obtained. The mass spectrum is a graph of relative intensity (related to the number of ions detected) versus mass-to-charge ratio. Mass spectrometry uses a very small sample size (10^{-12} g) to obtain results.

The resulting spectrum is like a puzzle. It contains numerous peaks that correspond to fragments of the initial molecule. The largest peak (parent peak) corresponds to the molecular mass of the molecular ion. By analyzing the peaks, scientists can determine the identity and structure of a compound. Computers are used to help interpret the spectrum and identify the molecule by using online spectral database libraries.

Mass spectrometry has been an essential tool for scientists since its invention in the early 1900s. But its use was limited to small molecules from which ion creation was easy. Large biological molecules could not be studied, because they would break down or decompose during conventional ion-formation techniques. In the late 1980s, two groups developed ion-formation methods that are used today in commercial mass spectrometers. John Fenn (Virginia Commonwealth University) developed electrospray ionization mass spectrometry, and Koichi Tanaka (Shimadzu Corporation, Japan) developed matrix-assisted laser desorption ionization (MALDI) mass spectrometry. In 2002, both scientists received the Nobel Prize in chemistry for their work. Their methods opened the field to the study of large molecules, allowing scientists to use mass spectrometry to study the structure of macromolecules, such as nucleic acids and steroids, and to identify the sequence of proteins.

Questions

1. Why is it necessary to convert the sample into ions in the mass spectrometer?
2. How have recent developments in mass spectrometry contributed to the field of medicine?



Scientists use mass spectrometers to identify and study the structure of molecules.

SECTION 3

Using Chemical Formulas

Key Terms
formula mass
percentage composition

Main Ideas

- Formula mass is the sum of the average atomic masses of a compound's elements.
- The molar mass of a compound is numerically equal to its formula mass.
- Molar mass is used to convert from moles to grams.
- Percent composition is the number of grams in one mole of a compound.

VIRGINIA STANDARDS

CH.3.c The student will investigate and understand how conservation of energy and matter is expressed in chemical formulas and balanced equations. Key concepts include: writing chemical formulas.
CH.1.EKS-19; CH.4.EKS-1

As you have seen, a chemical formula indicates the elements as well as the relative number of atoms or ions of each element present in a compound. Chemical formulas also allow chemists to calculate a number of characteristic values for a given compound. In this section, you will learn how to use chemical formulas to calculate the *formula mass*, the *molar mass*, and the *percentage composition* by mass of a compound.

MAIN IDEA

Formula mass is the sum of the average atomic masses of a compound's elements.

In an earlier chapter, we saw that hydrogen atoms have an average atomic mass of 1.007 94 u and that oxygen atoms have an average atomic mass of 15.9994 u. Like individual atoms, molecules, formula units, or ions have characteristic average masses. For example, we know from the chemical formula H_2O 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 adding 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 u
average atomic mass of O: 16.00 u

$$2\ \cancel{\text{H atoms}} \times \frac{1.01\ \text{u}}{\cancel{\text{H atom}}} = 2.02\ \text{u}$$

$$1\ \cancel{\text{O atom}} \times \frac{16.00\ \text{u}}{\cancel{\text{O atom}}} = 16.00\ \text{u}$$

$$\text{average mass of } H_2O \text{ molecule} = 18.02\ \text{u}$$

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 atoms represented in its formula.**

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 in the back of the book have been rounded to two decimal places.

Formula Mass

Sample Problem F **Find the formula mass of potassium chlorate, $KClO_3$.**

SOLVE

The mass of a formula unit of $KClO_3$ 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 in the back of the book. In the calculation, each atomic mass has been rounded to two decimal places.

$$1\ \cancel{\text{K atom}} \times \frac{39.10\ \text{u}}{\cancel{\text{K atom}}} = 39.10\ \text{u}$$

$$1\ \cancel{\text{Cl atom}} \times \frac{35.45\ \text{u}}{\cancel{\text{Cl atom}}} = 35.45\ \text{u}$$

$$3\ \cancel{\text{O atoms}} \times \frac{16.00\ \text{u}}{\cancel{\text{O atom}}} = 48.00\ \text{u}$$

$$\text{formula mass of } KClO_3 = 122.55\ \text{u}$$

Practice Answers in Appendix E

1. Find the formula mass of each of the following:
   a. H_2SO_4
   b. $Ca(NO_3)_2$
   c. PO_4^{3-}
   d. $MgCl_2$

PREMIUM CONTENT
Animated Chemistry
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Formula Mass and Molar Mass (Animation)

MAIN IDEA

The molar mass of a compound is numerically equal to its formula mass.

In an earlier chapter, you learned that the molar mass of a substance is equal to the mass in grams of one mole, or approximately 6.022×10^{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 adding 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.

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

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

$$\text{molar mass of } H_2O = 18.02\ \text{g/mol}$$

Figure 3.1 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 F the formula mass of $KClO_3$ was found to be 122.55 u. Therefore, because molar mass is numerically equal to formula mass, we know that the molar mass of $KClO_3$ is 122.55 g/mol.

FIGURE 3.1

Molar Mass Every compound has a characteristic molar mass. Shown here are one mole each of nitrogen (in balloon), water (in graduated cylinder), cadmium sulfide, CdS (yellow substance), and sodium chloride, NaCl (white substance).

Molar Mass

Sample Problem G **What is the molar mass of barium nitrate, $Ba(NO_3)_2$?**

SOLVE

One mole of barium nitrate contains exactly one mole of Ba^{2+} ions and two moles of NO_3^- ions. The two moles of NO_3^- ions contain two moles of N atoms and six moles of O atoms. Therefore, the molar mass of $Ba(NO_3)_2$ is calculated as follows.

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

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

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

$$\text{molar mass of } Ba(NO_3)_2 = 261.35\ \text{g/mol}$$

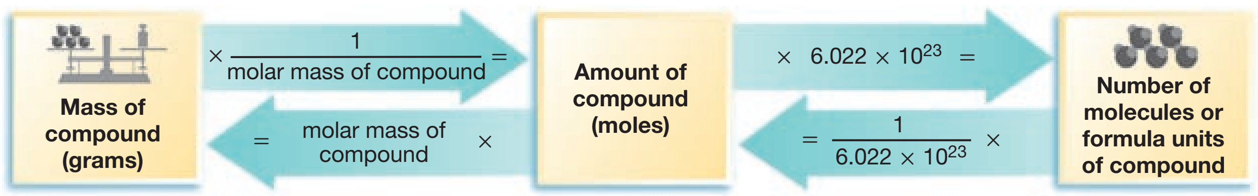
Practice

Answers in Appendix E

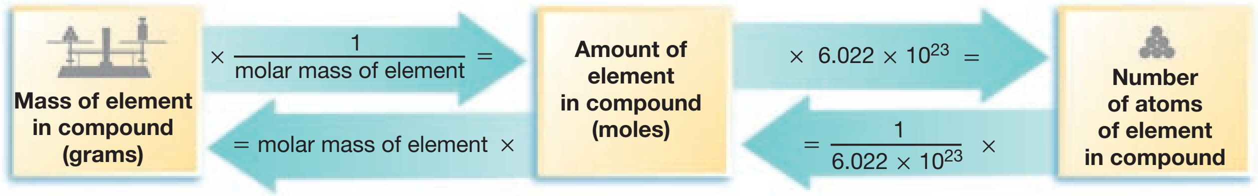
1. How many moles of atoms of each element are there in one mole of the following compounds?
   a. Al_2S_3
   b. $NaNO_3$
   c. $Ba(OH)_2$
2. Find the molar mass of each of the compounds listed in item 1.

FIGURE 3.2

Mass, Moles, and Molecules



(a) This 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.

MAIN IDEA

Molar mass is used to convert from moles to grams.

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 3.2.** Becoming familiar with this process now will help you a great deal, as mass-to-mole and mole-to-mass conversions will be widely used in future chapters.

Molar Mass as a Conversion Factor

Sample Problem H **What is the mass in grams of 2.50 mol of oxygen gas?**

1 ANALYZE

Given: 2.50 mol O_2

Unknown: mass of O_2 in grams

2 PLAN

moles $O_2 \longrightarrow$ grams O_2

To convert amount of O_2 in moles to mass of O_2 in grams, multiply by the molar mass of O_2.

$$\text{amount of } O_2 \text{ (mol)} \times \text{molar mass of } O_2 \text{ (g/mol)} = \text{mass of } O_2 \text{ (g)}$$

Continued

Molar Mass as a Conversion Factor (continued)

3 SOLVE

First, the molar mass of O_2 must be calculated.

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{\text{mol O}} = 32.00 \text{ g (mass of one mole of } O_2)$$

The molar mass of O_2 is therefore 32.00 g/mol. Now do the calculation shown in step 2.

$$2.50 \text{ mol } O_2 \times \frac{32.00 \text{ g } O_2}{\text{mol } O_2} = 80.0 \text{ g } O_2$$

4 CHECK YOUR WORK

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.

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

PREMIUM CONTENT

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Molar Mass as a Conversion Factor

Sample Problem I **Ibuprofen, $C_{13}H_{18}O_2$, is the active ingredient in many nonprescription pain relievers.**

Its molar mass is 206.31 g/mol.

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

1 ANALYZE

Given: 33 g of $C_{13}H_{18}O_2$, molar mass 206.31 g/mol

Unknown:
a. moles $C_{13}H_{18}O_2$
b. molecules $C_{13}H_{18}O_2$
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 $C_{13}H_{18}O_2$.

$$\text{g } C_{13}H_{18}O_2 \times \frac{1 \text{ mol } C_{13}H_{18}O_2}{206.31 \text{ g } C_{13}H_{18}O_2} = \text{mol } C_{13}H_{18}O_2$$

b. moles ⟶ molecules

To find the number of molecules of ibuprofen, multiply amount of $C_{13}H_{18}O_2$ in moles by Avogadro's number.

$$\text{mol } C_{13}H_{18}O_2 \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = \text{molecules } C_{13}H_{18}O_2$$

Continued

c. moles $C_{13}H_{18}O_2 \longrightarrow$ 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 $C_{13}H_{18}O_2$ and the molar mass of carbon.

$$\text{mol } C_{13}H_{18}O_2 \times \frac{13 \text{ mol C}}{\text{mol } C_{13}H_{18}O_2} \times \frac{12.01 \text{ g C}}{\text{mol C}} = \text{g C}$$

3 SOLVE

a. $33 \cancel{\text{g } C_{13}H_{18}O_2} \times \frac{1 \text{ mol } C_{13}H_{18}O_2}{206.31 \cancel{\text{g } C_{13}H_{18}O_2}} = 0.16 \text{ mol } C_{13}H_{18}O_2$

b. $0.16 \cancel{\text{mol}}\ C_{13}H_{18}O_2 \times \frac{6.022 \times 10^{23} \text{ molecules}}{\cancel{\text{mol}}} = 9.6 \times 10^{22} \text{ molecules } C_{13}H_{18}O_2$

c. $0.16 \cancel{\text{mol } C_{13}H_{18}O_2} \times \frac{13 \cancel{\text{mol C}}}{\cancel{\text{mol } C_{13}H_{18}O_2}} \times \frac{12.01 \text{ g C}}{\cancel{\text{mol C}}} = 25 \text{ g C}$

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

4 CHECK YOUR WORK

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

Practice

Answers in Appendix E

1. How many moles of compound are there in the following?
   a. 6.60 g $(NH_4)_2SO_4$
   b. 4.5 kg $Ca(OH)_2$
2. How many molecules are there in the following.
   a. 25.0 g H_2SO_4
   b. 125 g of sugar, $C_{12}H_{22}O_{11}$
3. What is the mass in grams of 6.25 mol of copper(II) nitrate?

MAIN IDEA

Percent composition is the number of grams in one mole of a compound.

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, $KClO_3$, 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, and then multiply this value by 100.

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

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.

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

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

Percentage Composition

Sample Problem J **Find the percentage composition of copper(I) sulfide, Cu_2S.**

1 ANALYZE

Given: formula, Cu_2S

Unknown: percentage composition of Cu_2S

2 PLAN

formula ⟶ molar mass ⟶ 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 SOLVE

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

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

$$\text{molar mass of } Cu_2S = 159.2\text{ g}$$

$$\frac{127.1\text{ g Cu}}{159.2\text{ g } Cu_2S} \times 100 = 79.85\%\text{ Cu}$$

$$\frac{32.07\text{ g S}}{159.2\text{ g } Cu_2S} \times 100 = 20.15\%\text{ S}$$

4 CHECK YOUR WORK

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%.)

Percentage Composition

Sample Problem K **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, $Na_2CO_3{\bullet}10H_2O$, which has a molar mass of 286.19 g/mol.**

Continued

Percentage Composition (continued)

1 ANALYZE

Given: chemical formula, $Na_2CO_3{\bullet}10H_2O$
molar mass of $Na_2CO_3{\bullet}10H_2O$

Unknown: mass percentage of H_2O

2 PLAN

chemical formula $\longrightarrow$ mass H_2O per mole of $Na_2CO_3{\bullet}10H_2O$ $\longrightarrow$ % 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 $Na_2CO_3{\bullet}10H_2O$.

3 SOLVE

One mole of $Na_2CO_3{\bullet}10H_2O$ contains 10 mol H_2O. As noted earlier, the molar mass of H_2O is 18.02 g/mol. The mass of 10 mol H_2O is calculated as follows.

$$10\ \cancel{\text{mol } H_2O} \times \frac{18.02\text{ g } H_2O}{\cancel{\text{mol } H_2O}} = 180.2\text{ g } H_2O$$

$$\text{mass of } H_2O \text{ per mole of } Na_2CO_3{\bullet}10H_2O = 180.2\text{ g}$$

The molar mass of $Na_2CO_3{\bullet}10H_2O$ is 286.19 g/mol, so we know that 1 mol of the hydrate has a mass of 286.19 g. The mass percentage of 10 mol H_2O in 1 mol $Na_2CO_3{\bullet}10H_2O$ can now be calculated.

$$\text{mass percentage of } H_2O \text{ in } Na_2CO_3{\bullet}10H_2O = \frac{180.2\text{ g } H_2O}{286.19\text{ g } Na_2CO_3{\bullet}10H_2O} \times 100$$

$$= 62.97\%\ H_2O$$

4 CHECK YOUR WORK

Checking shows that the arithmetic is correct and that units cancel as desired.

Practice Answers in Appendix E

1. Find the percentage compositions of the following:
   a. $PbCl_2$ **b.** $Ba(NO_3)_2$
2. Find the mass percentage of water in $ZnSO_4{\bullet}7H_2O$.
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?

SECTION 3 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. Determine both the formula mass and molar mass of ammonium carbonate, $(NH_4)_2CO_3$.
2. How many moles of atoms of each element are there in one mole of $(NH_4)_2CO_3$?
3. What is the mass in grams of 3.25 mol $Fe_2(SO_4)_3$?
4. How many molecules of aspirin, $C_9H_8O_4$, are there in a 100.0 mg tablet of aspirin?
5. Calculate the percentage composition of $(NH_4)_2CO_3$.

Critical Thinking

6. **RELATING IDEAS** A sample of hydrated copper (II) sulfate ($CuSO_4{\bullet}nH_2O$) is heated to 150°C and produces 103.74 g anhydrous copper (II) sulfate and 58.55 g water. How many moles of water molecules are present in 1.0 mol of hydrated copper (II) sulfate?

Determining Chemical Formulas

SECTION 4

Main Ideas

- Empirical formulas show the whole number ratio of elements in a compound.
- Molecular formulas give the types and numbers of atoms in a compound.

VIRGINIA STANDARDS

CH.3.c The student will investigate and understand how conservation of energy and matter is expressed in chemical formulas and balanced equations. Key concepts include: writing chemical formulas.

Key Terms

empirical formula

When a new substance is synthesized or is discovered, it is analyzed quantitatively to reveal its percentage composition. From these 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 BH_3, but the molecular formula is B_2H_6. In this case, the number of atoms given by the molecular formula corresponds to the empirical ratio multiplied by two.

MAIN IDEA

Empirical formulas show the whole number ratio of elements in a compound.

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

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

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

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 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 M.

Empirical Formulas

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

1 ANALYZE

Given: percentage composition: 32.38% Na, 22.65% S, and 44.99% O

Unknown: empirical formula

2 PLAN

percentage composition ⟶ mass composition ⟶ composition in moles ⟶ smallest whole-number mole ratio of atoms

3 SOLVE

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:

$$32.38 \cancel{\text{g Na}} \times \frac{1 \text{ mol Na}}{22.99 \cancel{\text{g Na}}} = 1.408 \text{ mol Na}$$

$$22.65 \cancel{\text{g S}} \times \frac{1 \text{ mol S}}{32.07 \cancel{\text{g S}}} = 0.7063 \text{ mol S}$$

$$44.99 \cancel{\text{g O}} \times \frac{1 \text{ mol O}}{16.00 \cancel{\text{g O}}} = 2.812 \text{ mol O}$$

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 Na_2SO_4.

4 CHECK YOUR WORK

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.

Empirical Formulas

PREMIUM CONTENT

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Sample Problem M **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?**

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 SOLVE

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

sample mass − phosphorus mass = 10.150 g − 4.433 g = 5.717 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 P_2O_5.

4 CHECK YOUR WORK

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

Answers in Appendix E

1. A compound is found to contain 63.52% iron and 36.48% sulfur. Find its empirical formula.
2. Find the empirical formula of a compound found to contain 26.56% potassium, 35.41% chromium, and the remainder oxygen.
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?

MAIN IDEA

Molecular formulas give the types and numbers of atoms in a compound.

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 BH_3. Any multiple of BH_3, such as B_2H_6, B_3H_9, B_4H_{12}, and so on, represents the same ratio of B atoms to H atoms. The molecular compounds ethene, C_2H_4, and cyclopropane, C_3H_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 u. The formula mass for the empirical formula, BH_3, is 13.84 u. Dividing the experimental formula mass by the empirical formula mass gives the value of x for diborane.

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

The molecular formula of diborane is therefore B_2H_6.

$$2(BH_3) = B_2H_6$$

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.

Molecular Formulas

Sample Problem N **In Sample Problem M, the empirical formula of a compound of phosphorus and oxygen was found to be P_2O_5. Experimentation shows that the molar mass of this compound is 283.89 g/mol. What is the compound's molecular formula?**

1 ANALYZE

Given: empirical formula

Unknown: molecular formula

Continued

Molecular Formulas (continued)

2 PLAN

x (empirical formula) = molecular formula

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

3 SOLVE

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

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

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

$$\text{mass of phosphorus atom} = 30.97 \text{ u}$$
$$\text{mass of oxygen atom} = 16.00 \text{ u}$$
$$\text{empirical formula mass of } P_2O_5 = 2 \times 30.97 \text{ u} + 5 \times 16.00 \text{ u} = 141.94 \text{ u}$$

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.

$$x = \frac{283.89 \cancel{\text{u}}}{141.94 \cancel{\text{u}}} = 2.0001$$

The compound's molecular formula is therefore P_4O_{10}.

$$2 \times (P_2O_5) = P_4O_{10}$$

4 CHECK YOUR WORK

Checking the arithmetic shows that it is correct.

Practice Answers in Appendix E

1. Determine the molecular formula of the compound with an empirical formula of CH and a formula mass of 78.110 u.
2. A sample of a compound with a formula mass of 34.00 u is found to consist of 0.44 g H and 6.92 g O. Find its molecular formula.

SECTION 4 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. A compound contains 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 u, what is the molecular formula of this compound?

Critical Thinking

5. **RELATING IDEAS** A compound containing sodium, chlorine, and oxygen is 25.42% sodium by mass. A 3.25 g sample gives 4.33×10^{22} atoms of oxygen. What is the empirical formula?

Math Tutor

Calculating Percentage Composition

Chemists can analyze an unknown substance by determining its percentage composition by mass. Percentage composition is determined by finding the mass of each element in a sample of the substance as a percentage of the mass of the whole sample. The results of this analysis can then be compared with the percentage composition of known compounds to determine the probable identity of the unknown substance. Once you know a compound's formula, you can determine its percentage composition by mass.

Sample Problem

Determine the percentage composition of potassium chlorate, $KClO_3$.

First, calculate the molar mass of $KClO_3$. The formula shows you that one mole of $KClO_3$ consists of 1 mol K atoms, 1 mol Cl atoms, and 3 mol O atoms. Thus, the molar mass of $KClO_3$ is molar mass K + molar mass Cl + 3(molar mass O) = 39.10 g K + 35.45 g Cl + 3(16.00 g O).

$$\text{molar mass } KClO_3 = 122.55 \text{ g}$$

The percentage composition of $KClO_3$ is determined by calculating the percentage of the total molar mass contributed by each element.

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

$$\% \text{ K in } KClO_3 = \frac{39.10 \text{ g K}}{122.55 \text{ g } KClO_3} \times 100 = 31.91\%$$

$$\% \text{ Cl in } KClO_3 = \frac{35.45 \text{ g Cl}}{122.55 \text{ g } KClO_3} \times 100 = 28.93\%$$

$$\% \text{ O in } KClO_3 = \frac{48.00 \text{ g O}}{122.55 \text{ g } KClO_3} \times 100 = 39.17\%$$

Determine the percentage of nitrogen in ammonium sulfate, $(NH_4)_2SO_4$.

Even though you want to find the percentage of only one element, you must calculate the molar mass of $(NH_4)_2SO_4$. To do that, examine the formula to find the number of moles of each element in the compound. The two ammonium groups, indicated by $(NH_4)_2$, contain 2 mol N and 8 mol H per mole of $(NH_4)_2SO_4$. The sulfate group, SO_4, contains 1 mol S and 4 mol O per mole of $(NH_4)_2SO_4$.

$$\begin{aligned}
2 \text{ mol N} = 2 \times 14.01 \text{ g} &= 28.02 \text{ g}\\
8 \text{ mol H} = 8 \times 1.01 \text{ g} &= 8.08 \text{ g}\\
1 \text{ mol S} = 1 \times 32.07 &= 32.07 \text{ g}\\
4 \text{ mol O} = 4 \times 16.00 &= \underline{64.00 \text{ g}}\\
\text{molar mass } (NH_4)_2SO_4 &= 132.17 \text{ g}
\end{aligned}$$

Now, you can determine the percentage of nitrogen in the compound as follows.

$$\% \text{ N in } (NH_4)_2SO_4 = \frac{28.02 \text{ g N}}{132.17 \text{ g } (NH_4)_2SO_4} \times 100 = 21.20\%$$

Practice

1. What is the percentage composition of sodium carbonate, Na_2CO_3?
2. What is the percentage of iodine in zinc iodate, $Zn(IO_3)_2$?

CHAPTER 7 Summary

PREMIUM CONTENT

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Review Games
Concept Maps

SECTION 1 Chemical Names and Formulas

- 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.
- Binary compounds are composed of two elements.
- Binary ionic compounds are named by combining the names of the positive and negative ions.
- The old system of naming binary molecular compounds uses prefixes. The new system, known as the *Stock system*, uses oxidation numbers.

KEY TERMS

monatomic ion
binary compound
nomenclature
oxyanion
salt

SECTION 2 Oxidation Numbers

- 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 by using the Stock system.
- Stock-system names and prefix-system names are used interchangeably for many molecular compounds.
- Oxidation numbers of each element in a 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.

KEY TERMS

oxidation number
oxidation state

SECTION 3 Using Chemical Formulas

- 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 is used as a conversion factor between amount in moles and mass in grams of a given compound or element.

KEY TERMS

formula mass
percentage composition

SECTION 4 Determining Chemical Formulas

- An empirical formula shows the simplest whole-number ratio of atoms in a given compound.
- 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.

KEY TERMS

empirical formula

CHAPTER 7 Review

SECTION 1

Chemical Names and Formulas

REVIEWING MAIN IDEAS

1. a. What are monatomic ions?
   b. Give three examples of monatomic ions.

2. How does the chemical formula for the nitrite ion differ from the chemical formula for the nitrate ion?

3. Using only the periodic table, write the symbol of the ion most typically formed by each of the following elements:
   a. K
   b. Ca
   c. S
   d. Cl
   e. Ba
   f. Br

4. Write the formula for and indicate the charge on each of the following ions:
   a. sodium ion
   b. aluminum ion
   c. chloride ion
   d. nitride ion
   e. iron(II) ion
   f. iron(III) ion

5. Name each of the following monatomic ions:
   a. K^+
   b. Mg^{2+}
   c. Al^{3+}
   d. Cl^-
   e. O^{2-}
   f. Ca^{2+}

6. Write formulas for the binary ionic compounds formed between the following elements. (Hint: See Sample Problem A.)
   a. sodium and iodine
   b. calcium and sulfur
   c. zinc and chlorine
   d. barium and fluorine
   e. lithium and oxygen

7. Give the name of each of the following binary ionic compounds. (Hint: See Sample Problem B.)
   a. KCl
   b. $CaBr_2$
   c. Li_2O
   d. $MgCl_2$

8. Write the formulas for and give the names of the compounds formed by the following ions:
   a. Cr^{2+} and F^-
   b. Ni^{2+} and O^{2-}
   c. Fe^{3+} and O^{2-}

9. What determines the order in which the component elements of binary molecular compounds are written?

10. Name the following binary molecular compounds according to the prefix system. (Hint: See Sample Problem D.)
    a. CO_2
    b. CCl_4
    c. PCl_5
    d. SeF_6
    e. As_2O_5

11. Write formulas for each of the following binary molecular compounds. (Hint: See Sample Problem D.)
    a. carbon tetrabromide
    b. silicon dioxide
    c. tetraphosphorus decoxide
    d. diarsenic trisulfide

12. Distinguish between binary acids and oxyacids, and give two examples of each.

13. a. What is a salt?
    b. Give two examples of salts.

14. Name each of the following acids:
    a. HF
    b. HBr
    c. HNO_3
    d. H_2SO_4
    e. H_3PO_4

15. Give the molecular formula for each of the following acids:
    a. sulfurous acid
    b. chloric acid
    c. hydrochloric acid
    d. hypochlorous acid
    e. perchloric acid
    f. carbonic acid
    g. acetic acid

PRACTICE PROBLEMS

16. 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

17. Name each of the following ions:

a. NH_4^+
b. ClO_3^-
c. OH^-
d. SO_4^{2-}
e. NO_3^-
f. CO_3^{2-}
g. PO_4^{3-}
h. CH_3COO^-
i. HCO_3^-
j. CrO_4^{2-}

18. 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

SECTION 2

Oxidation Numbers

REVIEWING MAIN IDEAS

19. Name each of the following ions according to the Stock system:

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

20. Name each of the binary molecular compounds in item 11 by using the Stock system.

21. Write formulas for each of the following compounds:

a. phosphorus(III) iodide
b. sulfur(II) chloride
c. carbon(IV) sulfide
d. nitrogen(V) oxide

22. **a.** What are oxidation numbers?
b. What useful functions do oxidation numbers serve?

PRACTICE PROBLEMS

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

a. NaCl
b. KF
c. CaS
d. $Co(NO_3)_2$
e. $FePO_4$
f. Hg_2SO_4
g. $Hg_3(PO_4)_2$

24. Assign oxidation numbers to each atom in the following compounds. (Hint: See Sample Problem E.)

a. HI
b. PBr_3
c. GeS_2
d. KH
e. As_2O_5
f. H_3PO_4

25. Assign oxidation numbers to each atom in the following ions. (Hint: See Sample Problem E.)

a. NO_3^-
b. ClO_4^-
c. PO_4^{3-}
d. $Cr_2O_7^{2-}$
e. CO_3^{2-}

SECTION 3

Using Chemical Formulas

REVIEWING MAIN IDEAS

26. **a.** Define *formula mass.*
b. In what unit is formula mass expressed?

27. What is meant by the molar mass of a compound?

PRACTICE PROBLEMS

28. Determine the formula mass of each of the following compounds or ions. (Hint: See Sample Problem F.)

a. glucose, $C_6H_{12}O_6$
b. calcium acetate, $Ca(CH_3COO)_2$
c. the ammonium ion, NH_4^+
d. the chlorate ion, ClO_3^-

29. 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. KNO_3
b. Na_2SO_4
c. $Ca(OH)_2$
d. $(NH_4)_2SO_3$
e. $Ca_3(PO_4)_2$
f. $Al_2(CrO_4)_3$

30. Determine the molar mass of each compound listed in item 29. (Hint: See Sample Problem G.)

31. Determine the number of moles of compound in each of the following samples. (Hint: See Sample Problem I.)
   a. 4.50 g H_2O
   b. 471.6 g $Ba(OH)_2$
   c. 129.68 g $Fe_3(PO_4)_2$

32. Determine the percentage composition of each of the following compounds. (Hint: See Sample Problem J.)
   a. NaCl
   b. $AgNO_3$
   c. $Mg(OH)_2$

33. Determine the percentage by mass of water in the hydrate $CuSO_4 \bullet 5H_2O$. (Hint: See Sample Problem K.)

SECTION 4

Determining Chemical Formulas

REVIEWING MAIN IDEAS

34. What three types of information are used to find an empirical formula from percentage composition data?

35. What is the relationship between the empirical formula and the molecular formula of a compound?

PRACTICE PROBLEMS

36. Determine the empirical formula of a compound containing 63.50% silver, 8.25% nitrogen, and 28.25% oxygen. (Hint: See Sample Problem L.)

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 CH_2O and a molar mass of 120.12 g/mol?

39. A compound with a formula mass of 42.08 u is found to be 85.64% carbon and 14.36% hydrogen by mass. Find its molecular formula.

Mixed Review

REVIEWING MAIN IDEAS

40. Chemical analysis shows that citric acid contains 37.51% C, 4.20% H, and 58.29% O. What is the empirical formula for citric acid?

41. Name each of the following compounds by using the Stock system:
   a. LiBr
   b. $Sn(NO_3)_2$
   c. $FeCl_2$
   d. MgO
   e. KOH
   f. Fe_2O_3
   g. $AgNO_3$
   h. $Fe(OH)_2$
   i. CrF_2

42. What is the mass in grams of each of the following samples?
   a. 1.000 mol NaCl
   b. 2.000 mol H_2O
   c. 3.500 mol $Ca(OH)_2$
   d. 0.625 mol $Ba(NO_3)_2$

43. Determine the formula mass and molar mass of each of the following compounds:
   a. XeF_4
   b. $C_{12}H_{24}O_6$
   c. Hg_2I_2
   d. 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, $Fe(CHO_2)_3 \bullet H_2O$? 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. HNO_2
   b. H_2SO_3
   c. H_2CO_3
   d. HI

47. Determine the percentage composition of the following compounds:
   a. NaClO
   b. H_2SO_3
   c. C_2H_5COOH
   d. $BeCl_2$

48. Name each of the following binary compounds:
   a. MgI_2
   b. NaF
   c. CS_2
   d. N_2O_4
   e. SO_2
   f. PBr_3
   g. $CaCl_2$
   h. AgI

49. Assign oxidation numbers to each atom in the following molecules and ions:
   a. CO_2
   b. NH_4^+
   c. MnO_4^-
   d. $S_2O_3^{2-}$
   e. H_2O_2
   f. P_4O_{10}
   g. OF_2

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 the compound's 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. Analyze the formula for sulfur trioxide. Then, use your analysis to list all of the chemical information that you can.

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?

USING THE HANDBOOK

53. Review the common reactions of Group 1 metals in the *Elements Handbook* (Appendix A), and answer the following questions:
   a. Some of the Group 1 metals react with oxygen to form superoxides. Write the formulas for these compounds.
   b. What is the charge on each cation for the formulas that you wrote in (a)?
   c. How does the charge on the anion vary for oxides, peroxides, and superoxides?

54. Review the common reactions of Group 2 metals in the *Elements Handbook* (Appendix A), and answer the following questions:
   a. Some of the Group 2 metals react with oxygen to form oxides. Write the formulas for these compounds.
   b. Some of the Group 2 metals react with oxygen to form peroxides. Write the formulas for these compounds.
   c. Some of the Group 2 metals react with nitrogen to form nitrides. Write the formulas for these compounds.
   d. Most Group 2 elements form hydrides. What is hydrogen's oxidation state in these compounds?

55. Review the analytical tests for transition metals in the *Elements Handbook* (Appendix A), and answer the following questions:
   a. Determine the oxidation state of each metal in the precipitates shown for cadmium, zinc, and lead.
   b. Determine the oxidation state of each metal in the complex ions shown for iron, manganese, and cobalt.
   c. The copper compound shown is called a *coordination compound*. The ammonia shown in the formula exists as molecules that do not have a charge. Determine copper's oxidation state in this compound.

56. Review the common reactions of Group 15 elements in the *Elements Handbook* (Appendix A), and answer the following questions:
   a. Write formulas for each of the oxides listed for the Group 15 elements.
   b. Determine nitrogen's oxidation state in the oxides listed in (a).

RESEARCH AND WRITING

57. **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 *Linnaean system of binomial nomenclature*, 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.

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

ALTERNATIVE ASSESSMENT

59. **Performance Assessment** Your teacher will supply you with a note card that has one of the following formulas on it: $NaCH_3COO \bullet 3H_2O$, $MgCl_2 \bullet 6H_2O$, $LiC_2H_3O_2 \bullet 2H_2O$, or $MgSO_4 \bullet 7H_2O$. 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?

60. Both ammonia, NH_3, and ammonium nitrate, NH_4NO_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 compound 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, and shipping costs.)

Standards-Based Assessment

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Which of the following compounds does not contain a polyatomic ion?
A. sodium carbonate
B. sodium sulfate
C. sodium sulfite
D. sodium sulfide

2. The correct formula for ammonium phosphate is
A. $(NH_4)_3PO_4$.
B. $(NH_4)_2PO_4$.
C. NH_4PO_4.
D. $NH_4(PO_4)_2$.

3. When writing the formula for a compound that contains a polyatomic ion,
A. write the anion's formula first.
B. use superscripts to show the number of polyatomic ions present.
C. use parentheses if the number of polyatomic ions is greater than 1.
D. always place the polyatomic ion in parentheses.

4. The correct name for NH_4CH_3COO is
A. ammonium carbonate.
B. ammonium hydroxide.
C. ammonium acetate.
D. ammonium nitrate.

5. Which of the following is the correct formula for iron (III) sulfate?
A. Fe_3SO_4
B. $Fe_3(SO_4)_2$
C. $Fe_2(SO_4)_3$
D. $3FeSO_4$

6. The molecular formula for acetylene is C_2H_2. The molecular formula for benzene is C_6H_6. The empirical formula for both is
A. CH.
B. C_2H_2.
C. C_6H_6.
D. $(CH)_2$.

7. Which of the following shows the percentage composition of H_2SO_4?
A. 2.5% H, 39.1% S, 58.5% O
B. 2.1% H, 32.7% S, 65.2% O
C. 28.6% H, 14.3% S, 57.1% O
D. 33.3% H, 16.7% S, 50% O

8. Which of the following compounds has the highest percentage of oxygen?
A. CH_4O
B. CO_2
C. H_2O
D. Na_2CO_3

9. The empirical formula for a compound that is 1.2% H, 42.0% Cl, and 56.8% O is
A. HClO.
B. $HClO_2$.
C. $HClO_3$.
D. $HClO_4$.

SHORT ANSWER

10. When a new substance is synthesized or is discovered experimentally, the substance is analyzed quantitatively. What information is obtained from this typical analysis, and how is this information used?

11. An oxide of selenium is 28.8% O. Find the empirical formula. Assuming that the empirical formula is also the molecular formula, name the oxide.

EXTENDED RESPONSE

12. What is an empirical formula, and how does it differ from a molecular formula?

13. What are Stock system names based on?

Test Tip
Whenever possible, highlight or underline numbers or words critical to answering a question correctly.