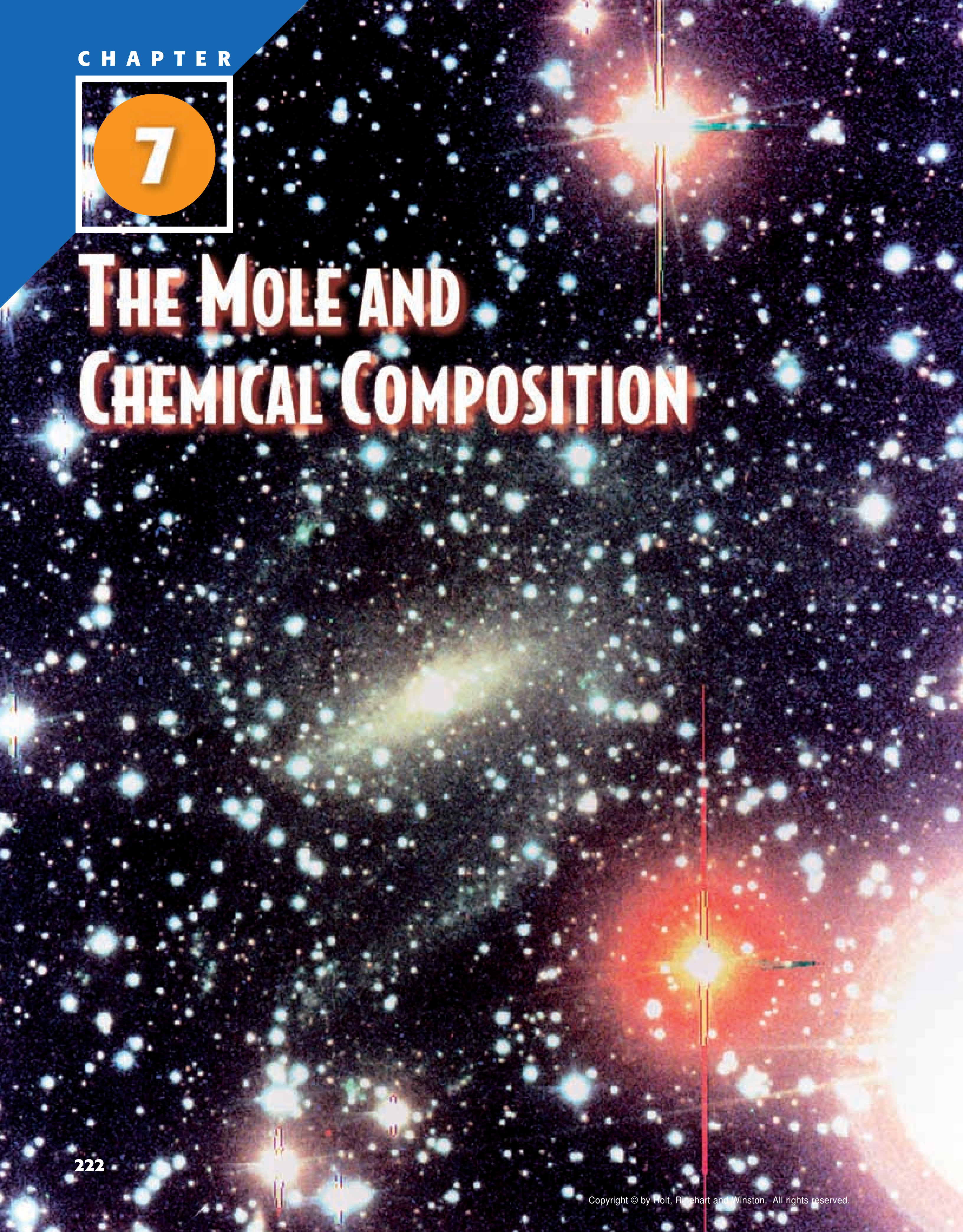
CHAPTER 7

The Mole and Chemical Composition



Galaxies have hundreds of billions of stars. The universe may have as many as sextillion stars—that's 1000 000 000 000 000 000 000 (or 1×10^{21}) stars. Such a number is called *astronomical* because it is so large that it usually refers only to vast quantities such as those described in astronomy. Can such a large number describe quantities that are a little more down to Earth? It certainly can. In fact, it takes an even larger number to describe the number of water molecules in a glass of water! In this chapter, you will learn about the *mole*, a unit used in chemistry to make working with such large quantities a little easier.

START-UP ACTIVITY

Counting Large Numbers

PROCEDURE

1. Count out exactly **200 small beads.** Using a **stopwatch,** record the amount of time it takes you to count them.
2. Your teacher will tell you the approximate number of small beads in 1 g. Knowing that number, calculate the mass of 200 small beads. Record the mass that you have calculated.
3. Use a **balance** to determine the mass of the 200 small beads that you counted in step 1. Compare this mass with the mass you calculated in step 2.
4. Using the mass you calculated in step 2 and a balance, measure out **another 200 small beads.** Record the amount of time it takes you to count small beads when using this counting method.
5. Count the number of **large beads** in 1 g.

ANALYSIS

1. Which method of counting took the most time?
2. Which method of counting do you think is the most accurate?
3. In a given mass, how does the number of large beads compare with the number of small beads? Explain your results.

Pre-Reading Questions

1. **What are some things that are sold by weight instead of by number?**
2. **Which would need a larger package, a kilogram of pencils or a kilogram of drinking straws?**
3. **If you counted one person per second, how many hours would it take to count the 6 billion people now in the world?**

CONTENTS 7

SECTION 1

Avogadro's Number and Molar Conversions

KEY TERMS

- **mole**
- **Avogadro's number**
- **molar mass**

OBJECTIVES

1. **Identify** the mole as the unit used to count particles, whether atoms, ions, or molecules.
2. **Use** Avogadro's number to convert between amount in moles and number of particles.
3. **Solve** problems converting between mass, amount in moles, and number of particles using Avogadro's number and molar mass.

Avogadro's Number and the Mole

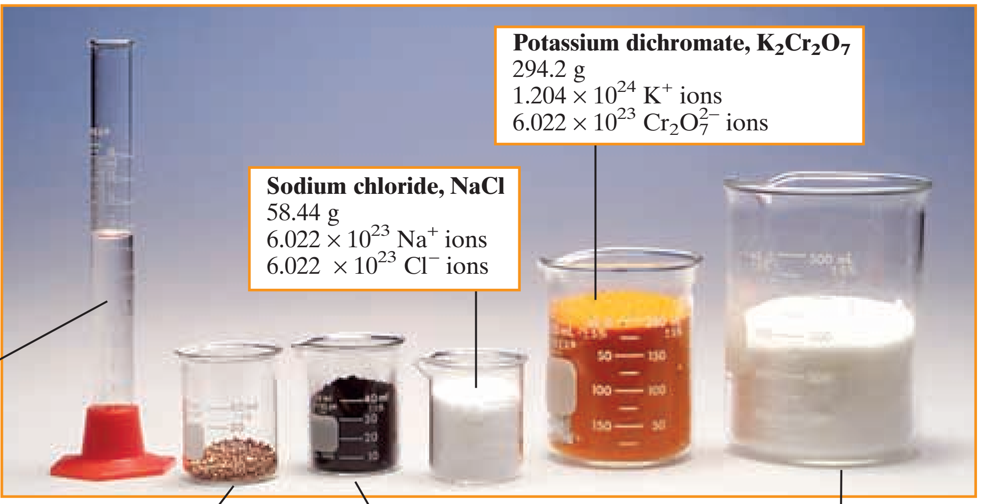
mole
the SI base unit used to measure the amount of a substance whose number of particles is the same as the number of atoms of carbon in exactly 12 grams of carbon-12

Avogadro's number
6.022×10^{23}, the number of atoms or molecules in 1.000 mol

Atoms, ions, and molecules are very small, so even tiny samples have a huge number of particles. To make counting such large numbers easier, scientists use the same approach to represent the number of ions or molecules in a sample as they use for atoms. The SI unit for amount is called the **mole** (mol). A mole is the number of atoms in exactly 12 grams of carbon-12.

The number of particles in a mole is called **Avogadro's number,** or Avogadro's constant. One way to determine this number is to count the number of particles in a small sample and then use mass or particle size to find the amount in a larger sample. This method works only if all of the atoms in the sample are identical. Thus, scientists measure Avogadro's number using a sample that has atoms of only one isotope.

MOLAR QUANTITIES OF SOME SUBSTANCES



Potassium dichromate, $K_2Cr_2O_7$
294.2 g
1.204×10^{24} K^+ ions
6.022×10^{23} $Cr_2O_7^{2-}$ ions

Sodium chloride, NaCl
58.44 g
6.022×10^{23} Na^+ ions
6.022×10^{23} Cl^- ions

Water, H_2O
18.02 g
6.022×10^{23} molecules

Copper, Cu
63.55 g
6.022×10^{23} atoms

Carbon, C
12.01 g
6.022×10^{23} atoms

Sucrose, $C_{12}H_{22}O_{11}$
342.34 g
6.022×10^{23} molecules

Figure 1 The particles in a mole can be atoms, molecules, or ions. Examples of a variety of molar quantities are given. Notice that the volume and mass of a molar quantity varies from substance to substance.

Table 1 Counting Units

Unit	Example
1 dozen	12 objects
1 score	20 objects
1 roll	50 pennies
1 gross	144 objects
1 ream	500 sheets of paper
1 hour	3600 seconds
1 mole	6.022×10^{23} particles

Figure 2
You can use mass to count out a roll of new pennies; 50 pennies are in a roll. One roll weighs about 125 g.

The most recent measurement of Avogadro's number shows that it is $6.02214199 \times 10^{23}$ units/mole. In this book, the measurement is rounded to 6.022×10^{23} units/mol. Avogadro's number is used to count any kind of particle, as shown in **Figure 1.**

The Mole Is a Counting Unit

Keep in mind that the mole is used to count out a given number of particles, whether they are atoms, molecules, formula units, ions, or electrons. The mole is used in the same way that other, more familiar counting units, such as those in **Table 1,** are used. For example, there are 12 eggs in one dozen eggs. You might want to know how many eggs are in 15 dozen. You can calculate the number of eggs by using a conversion factor as follows.

$$15 \cancel{\text{dozen eggs}} \times \frac{12 \text{ eggs}}{1 \cancel{\text{dozen eggs}}} = 180 \text{ eggs}$$

Figure 2 shows another way that you can count objects: by using mass.

Quick LAB

SAFETY PRECAUTIONS

Exploring the Mole

PROCEDURE

1. Use a **periodic table** to find the atomic mass of the following substances: **graphite (carbon), iron filings, sulfur powder, aluminum foil,** and **copper wire.**
2. Use a **balance** to measure out 1 mol of each substance.
3. Use **graduated beakers** to find the approximate volume in 1 mol of each substance.

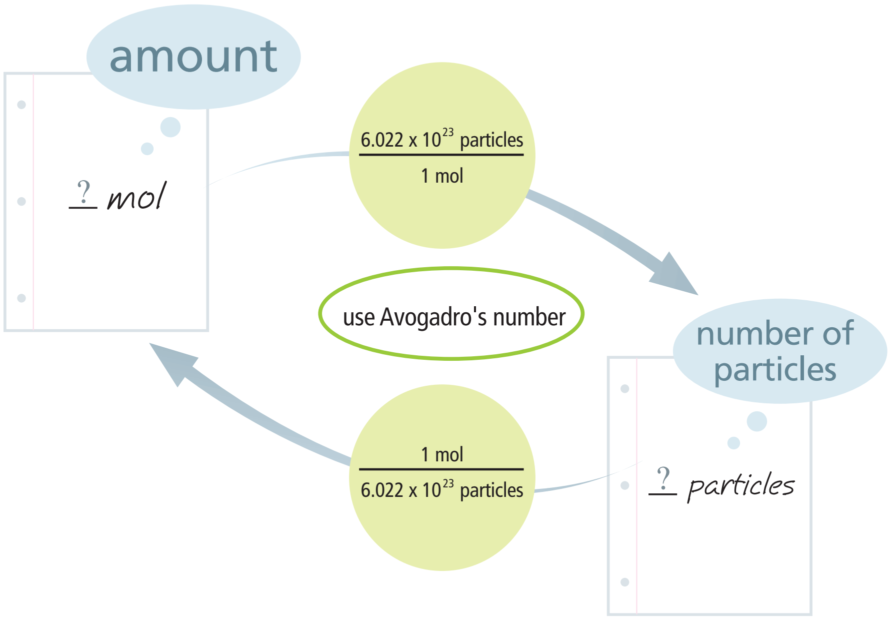
ANALYSIS

1. Which substance has the greatest atomic mass?
2. Which substance has the greatest mass in 1 mol?
3. Which substance has the greatest volume in 1 mol?
4. Does the mass of a mole of a substance relate to the substance's atomic mass?
5. Does the volume of a mole of a substance relate to the substance's atomic mass?

1 SKILLS Toolkit

Converting Between Amount in Moles and Number of Particles

1. Decide which quantity you are given: amount (in moles) or number of particles (in atoms, molecules, formula units, or ions).
2. If you are converting from amount to number of particles (going left to right), use the top conversion factor.
3. If you are converting from number of particles to amount (going right to left), use the bottom conversion factor.



Amount in Moles Can Be Converted to Number of Particles

A conversion factor begins with a definition of a relationship. The definition of one mole is

$$6.022 \times 10^{23} \text{ particles} = 1 \text{ mol}$$

If two quantities are equal and you divide one by the other, the factor you get is equal to 1. The following equation shows how this relationship is true for the definition of the mole.

$$\frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mol}} = 1$$

The factor on the left side of the equation is a conversion factor. The reciprocal of a conversion factor is also a conversion factor and is also equal to one, so the following is true.

$$\frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mol}} = \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ particles}} = 1$$

Because a conversion factor is equal to 1, it can multiply any quantity without changing the quantity's value. Only the units are changed.

These conversion factors can be used to convert between a number of moles of substance and a corresponding number of molecules. For example, imagine that you want to convert 2.66 mol of a compound into the corresponding number of molecules. How do you know which conversion factor to use? **Skills Toolkit 1** can help.

Choose the Conversion Factor That Cancels the Given Units

Take the amount (in moles) that you are given, shown in **Skills Toolkit 1** on the left, and multiply it by the conversion factor, shown in the top green circle, to get the number of particles, shown on the right. The calculation is as follows:

$$2.66 \cancel{\text{mol}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \cancel{\text{mol}}} = 1.60 \times 10^{24} \text{ molecules}$$

You can tell which of the two conversion factors to use, because the needed conversion factor should cancel the units of the given quantity to give you the units of the answer or the unknown quantity.

SKILLS Toolkit 2

PROBLEM SOLVING SKILL

Working Practice Problems

1. Gather information.

- Read the problem carefully.
- List the quantities and units given in the problem.
- Determine what value is being asked for (the answer) and the units it will need.

2. Plan your work.

Write the value of the given quantity times a question mark (which stands for a conversion factor) and then the equals sign, followed by another question mark (which stands for the answer) and the units of the answer. For example:

$$4.2 \text{ mol } CO_2 \times ? = ? \text{ molecules } CO_2$$

3. Calculate.

- Determine the conversion factor(s) needed to change the units of the given quantity to the units of the answer. Write the conversion factor(s) in the order you need them to cancel units.
- Cancel units, and check that the units that remain are the same on both sides and are the units desired for the answer.
- Calculate and round off the answer to the correct number of significant figures.
- Report your answer with correct units.

4. Verify your result.

- Verify your answer by estimating. One way to do so is to round off the numbers in the setup and make a quick calculation.
- Make sure your answer is reasonable. For example, if the number of atoms is less than one, the answer cannot possibly be correct.

STUDY TIP

Working Problems

If you have difficulty working practice problems, review the outline of procedures in **Skills Toolkit 2.** You may also refer back to the sample problems.

SAMPLE PROBLEM A

Converting Amount in Moles to Number of Particles

Find the number of molecules in 2.5 mol of sulfur dioxide.

1 Gather information.

- amount of SO_2 = 2.5 mol
- 1 mol of any substance = 6.022×10^{23} particles
- number of molecules of SO_2 = ? molecules

2 Plan your work.

The setup is: 2.5 mol $SO_2 \times ? = ?$ molecules SO_2

3 Calculate.

You are converting from the unit *mol* to the unit *molecules*. The conversion factor must have the units of *molecules/mol*. **Skills Toolkit 1** shows that this means you use 6.022×10^{23} molecules/1 mol.

$$2.5\ \cancel{\text{mol SO}_2} \times \frac{6.022 \times 10^{23}\ \text{molecules SO}_2}{1\ \cancel{\text{mol SO}_2}} = 1.5 \times 10^{24}\ \text{molecules SO}_2$$

4 Verify your result.

The units cancel correctly. The answer is greater than Avogadro's number, as expected, and has two significant figures.

PRACTICE HINT

Take your time, and be systematic. Focus on units; if they are not correct, you must rethink your preliminary equation. In this way, you can prevent mistakes.

PRACTICE

PROBLEM SOLVING SKILL

1. How many ions are there in 0.187 mol of Na^+ ions?
2. How many atoms are there in 1.45×10^{-17} mol of arsenic?
3. How many molecules are there in 4.224 mol of acetic acid, $C_2H_4O_2$?
4. How many formula units are there in 5.9 mol of NaOH?

Number of Particles Can Be Converted to Amount in Moles

Notice in **Skills Toolkit 1** that the reverse calculation is similar but that the conversion factor is inverted to get the correct units in the answer. Look at the following problem. How many moles are 2.54×10^{22} iron(III) ions, Fe^{3+}?

$$2.54 \times 10^{22}\ \text{ions Fe}^{3+} \times ? = ?\ \text{mol Fe}^{3+}$$

Multiply by the conversion factor that cancels the unit of *ions* and leaves the unit of *mol*. (That is, you use the conversion factor that has the units that you want to get on top and the units that you want to get rid of on the bottom.)

$$2.54 \times 10^{22}\ \cancel{\text{ions Fe}^{3+}} \times \frac{1\ \text{mol Fe}^{3+}}{6.022 \times 10^{23}\ \cancel{\text{ions Fe}^{3+}}} = 0.0422\ \text{mol Fe}^{3+}$$

This answer makes sense, because you started with fewer than Avogadro's number of ions, so you have less than one mole of ions.

SAMPLE PROBLEM B

Converting Number of Particles to Amount in Moles

A sample contains 3.01×10^{23} molecules of sulfur dioxide, SO_2. Determine the amount in moles.

1 Gather information.

- number of molecules of $SO_2 = 3.01 \times 10^{23}$ molecules
- 1 mol of any substance $= 6.022 \times 10^{23}$ particles
- amount of SO_2 = ? mol

2 Plan your work.

The setup is similar to the calculation in **Sample Problem A.**

$$3.01 \times 10^{23} \text{ molecules } SO_2 \times ? = ? \text{ mol } SO_2$$

3 Calculate.

The conversion factor is used to remove the unit of *molecules* and introduce the unit of *mol.*

$$3.01 \times 10^{23} \cancel{\text{molecules } SO_2} \times \frac{1 \text{ mol } SO_2}{6.022 \times 10^{23} \cancel{\text{molecules } SO_2}} = 0.500 \text{ mol } SO_2$$

4 Verify your result.

There are fewer than 6.022×10^{23} (Avogadro's number) of SO_2 molecules, so it makes sense that the result is less than 1 mol. Three is the correct number of significant figures.

PRACTICE HINT

Always check your answer for the correct number of significant figures.

PRACTICE

PROBLEM SOLVING SKILL

1. How many moles of xenon do 5.66×10^{23} atoms equal?
2. How many moles of silver nitrate do 2.888×10^{15} formula units equal?
3. A biologist estimates that there are 2.7×10^{17} termites on Earth. How many moles of termites is this?
4. How many moles do 5.66×10^{25} lithium ions, Li^+, equal?
5. Determine the number of moles of each specified atom or ion in the given samples of the following compounds. (Hint: The formula tells you how many atoms or ions are in each molecule or formula unit.)
   - **a.** O atoms in 3.161×10^{21} molecules of CO_2
   - **b.** C atoms in 3.161×10^{21} molecules of CO_2
   - **c.** O atoms in 2.222×10^{24} molecules of NO
   - **d.** K^+ ions in 5.324×10^{16} formula units of KNO_2
   - **e.** Cl^- ions in 1.000×10^{14} formula units of $MgCl_2$
   - **f.** N atoms in 2.000×10^{14} formula units of $Ca(NO_3)_2$
   - **g.** O atoms in 4.999×10^{25} formula units of $Mg_3(PO_4)_2$

Molar Mass Relates Moles to Grams

In chemistry, you often need to know the mass of a given number of moles of a substance or the number of moles in a given mass. Fortunately, the mole is defined in a way that makes figuring out either of these easy.

Amount in Moles Can Be Converted to Mass

molar mass

the mass in grams of one mole of a substance

The mole is the SI unit for amount. The **molar mass,** or mass in grams of one mole of an element or compound, is numerically equal to the atomic mass of monatomic elements and the formula mass of compounds and diatomic elements. To find a monatomic element's molar mass, use the atomic mass, but instead of having units of *amu,* the molar mass will have units of *g/mol.* So, the molar mass of carbon is 12.01 g/mol, and the molar mass of iron is 55.85 g/mol. How to find the molar mass of compounds and diatomic elements is shown in the next section.

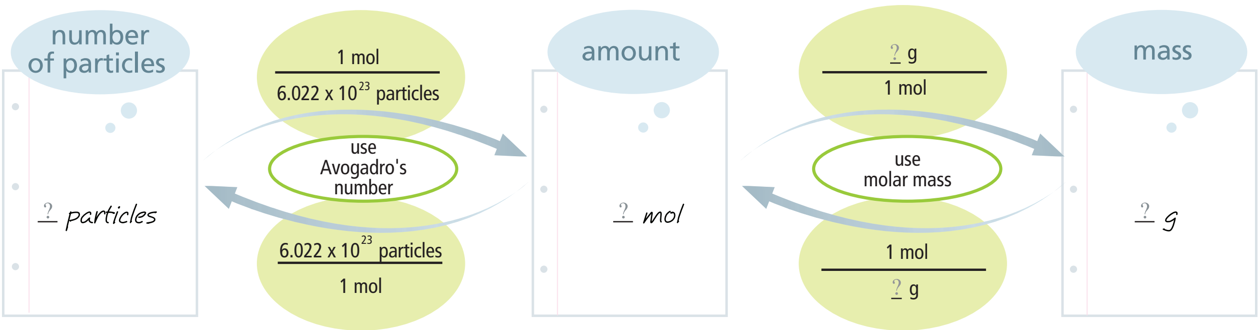
You use molar masses as conversion factors in the same way you use Avogadro's number. The right side of **Skills Toolkit 3** shows how the *amount* in moles relates to the *mass* in grams of a substance. Suppose you must find the mass of 3.50 mol of copper. You will use the molar mass of copper. By checking the periodic table, you find the atomic mass of copper, 63.546 amu, which you round to 63.55 amu. So, in calculations with copper, use 63.55 g/mol.

The Mole Plays a Central Part in Chemical Conversions

You know how to convert from number of particles to amount in moles and how to convert from amount in moles to mass. Now you can use the same methods one after another to convert from *number of particles* to *mass.* **Skills Toolkit 3** shows the two-part process for this conversion. One step common to many problems in chemistry is converting to amount in moles. **Sample Problem C** shows how to convert from number of particles to the mass of a substance by first converting to amount in moles.

3 SKILLS Toolkit

Converting Between Mass, Amount, and Number of Particles



SAMPLE PROBLEM C

Converting Number of Particles to Mass

Find the mass in grams of 2.44×10^{24} atoms of carbon, whose molar mass is 12.01 g/mol.

1 Gather information.

- number of atoms C = 2.44×10^{24} atoms
- molar mass of carbon = 12.01 g/mol
- amount of C = ? mol
- mass of the sample of carbon = ? g

2 Plan your work.

- **Skills Toolkit 3** shows that to convert from number of atoms to mass in grams, you must first convert to amount in moles.
- To find the amount in moles, select the conversion factor that will take you from number of atoms to amount in moles.

$$2.44 \times 10^{24} \text{ atoms} \times ? = ? \text{ mol}$$

- Multiply the number of atoms by the following conversion factor:

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

- To find the mass in grams, select the conversion factor that will take you from amount in moles to mass in grams.

$$? \text{ mol} \times ? = ? \text{ g}$$

- Multiply the amount in moles by the following conversion factor:

$$\frac{12.01 \text{ g C}}{1 \text{ mol}}$$

3 Calculate.

Solve and cancel identical units in the numerator and denominator.

$$2.44 \times 10^{24} \cancel{\text{atoms}} \times \frac{1 \cancel{\text{mol}}}{6.022 \times 10^{23} \cancel{\text{atoms}}} \times \frac{12.01 \text{ g C}}{1 \cancel{\text{mol}}} = 48.7 \text{ g C}$$

4 Verify your result.

The answer has the units requested in the problem.

PRACTICE HINT

Make sure to select the correct conversion factors so that units cancel to get the unit required in the answer.

PRACTICE

Given molar mass, find the mass in grams of each of the following substances:

1. 2.11×10^{24} atoms of copper (molar mass of Cu = 63.55 g/mol)
2. 3.01×10^{23} formula units of NaCl (molar mass of NaCl = 58.44 g/mol)
3. 3.990×10^{25} molecules of CH_4 (molar mass of CH_4 = 16.05 g/mol)
4. 4.96 mol titanium (molar mass of Ti = 47.88 g/mol)

PROBLEM SOLVING SKILL

Mass Can Be Converted to Amount in Moles

Converting from mass to number of particles is the reverse of the operation in the previous problem. This conversion is also shown in **Skills Toolkit 3,** but this time you are going from right to left and using the bottom conversion factors.

Sample Problem D shows how to convert the mass of a substance to amount (mol) and then convert amount to the number of particles. Notice that the problem is the reverse of **Sample Problem C.**

PRACTICE HINT

Because no elements have a molar mass less than one, the number of grams in a sample of a substance will always be larger than the number of moles of the substance. Thus, when you convert from grams to moles, you will get a smaller number. And the opposite is true for the reverse calculation.

SAMPLE PROBLEM D

Converting Mass to Number of Particles

Find the number of molecules present in 47.5 g of glycerol, $C_3H_8O_3$. The molar mass of glycerol is 92.11 g/mol.

1 Gather information.

- mass of the sample of $C_3H_8O_3$ = 47.5 g
- molar mass of $C_3H_8O_3$ = 92.11 g/mol
- amount of $C_3H_8O_3$ = ? mol
- number of molecules $C_3H_8O_3$ = ? molecules

2 Plan your work.

- **Skills Toolkit 3** shows that you must first find the amount in moles.
- To determine the amount in moles, select the conversion factor that will take you from mass in grams to amount in moles.

$$47.5 \text{ g} \times ? = ? \text{ mol}$$

- Multiply mass by the conversion factor $\dfrac{1 \text{ mol}}{92.11 \text{ g } C_3H_8O_3}$
- To determine the number of particles, select the conversion factor that will take you from amount in moles to number of particles.

$$? \text{ mol} \times ? = ? \text{ molecules}$$

- Multiply amount by the conversion factor $\dfrac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}}$

3 Calculate.

$$47.5 \cancel{\text{ g } C_3H_8O_3} \times \frac{1 \cancel{\text{mol}}}{92.11 \cancel{\text{ g } C_3H_8O_3}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \cancel{\text{mol}}} = 3.11 \times 10^{23} \text{ molecules}$$

4 Verify your result.

The answer has the units requested in the problem.

PRACTICE

1. Find the number of atoms in 237 g Cu (molar mass of Cu = 63.55 g/mol).
2. Find the number of ions in 20.0 g Ca^{2+} (molar mass of Ca^{2+} = 40.08 g/mol).
3. Find the number of atoms in 155 mol of arsenic.

PROBLEM SOLVING SKILL

1 H Hydrogen 1.007 94 $1s^1$	11 Na Sodium 22.989 770 $[Ne]3s^1$	17 Cl Chlorine 35.4527 $[Ne]3s^23p^5$
1.01 g/mol	22.99 g/mol	35.45 g/mol

Figure 3
Round molar masses from the periodic table to two significant figures to the right of the decimal point.

Remember to Round Consistently

Calculators may report many figures. However, an answer must never be given to more figures than is appropriate. If the given amount has only two significant figures, then you must round the calculated number off to two significant figures. Also, keep in mind that many numbers are exact. In the definition of the mole, the chosen amount is *exactly* 12 grams of the carbon-12 isotope. Such numbers are not considered when rounding. **Figure 3** shows how atomic masses are rounded in this text.

1 Section Review

UNDERSTANDING KEY IDEAS

1. What is the definition of a mole?

2. How many particles are there in one mole?

3. Explain how Avogadro's number can give two conversion factors.

4. Which will have the greater number of ions, 1 mol of nickel(II) or 1 mol of copper(I)?

5. Without making a calculation, is 1.11 mol Pt more or less than 6.022×10^{23} atoms?

PRACTICE PROBLEMS

6. Find the number of molecules or ions.

a. 2.00 mol Fe^{3+} **c.** 0.25 mol K^+

b. 4.5 mol BCl_3 **d.** 6.022 mol O_2

7. Find the number of sodium ions, Na^+.

a. 3.00 mol Na_2CO_3

b. 3.00 mol $Na_4P_2O_7$

c. 5.12 mol $NaNO_3$

8. Find the number of moles.

a. 3.01×10^{23} molecules H_2O

b. 1.000×10^{23} atoms C

c. 5.610×10^{22} ions Na^+

9. Find the mass in grams.

a. 4.30×10^{16} atoms He, 4.00 g/mol

b. 5.710×10^{23} molecules CH_4, 16.05 g/mol

c. 3.012×10^{24} ions Ca^{2+}, 40.08 g/mol

10. Find the number of molecules or ions.

a. 1.000 g I^-, 126.9 g/mol

b. 3.5 g Cu^{2+}, 63.55 g/mol

c. 4.22 g SO_2, 64.07 g/mol

11. What is the mass of 6.022×10^{23} molecules of ibuprofen (molar mass of 206.31 g/mol)?

12. Find the mass in grams.

a. 4.01×10^{23} atoms Ca, 40.08 g/mol

b. 4.5 mol boron-11, 11.01 g/mol

c. 1.842×10^{19} ions Na^+, 22.99 g/mol

13. Find the number of molecules.

a. 2.000 mol H_2, 2.02 g/mol

b. 4.01 g HF, 20.01 g/mol

c. 4.5 mol $C_6H_{12}O_6$, 180.18 g/mol

CRITICAL THINKING

14. Why do we use carbon-12 rather than ordinary carbon as the basis for the mole?

15. Use **Skills Toolkit 1** to explain how a number of atoms is converted into amount in moles.

SECTION 2

Relative Atomic Mass and Chemical Formulas

KEY TERM

- **average atomic mass**

OBJECTIVES

1. **Use** a periodic table or isotopic composition data to determine the average atomic masses of elements.
2. **Infer** information about a compound from its chemical formula.
3. **Determine** the molar mass of a compound from its formula.

Average Atomic Mass and the Periodic Table

Topic Link

Refer to the "Atoms and Moles" chapter for a discussion of atomic mass and isotopes.

You have learned that you can use atomic masses on the periodic table to find the molar mass of elements. Many of these values on the periodic table are close to whole numbers. However, most atomic masses are written to at least three places past the decimal.

Why are the atomic masses of most elements on the periodic table not exact whole numbers? One reason is that the masses reported are *relative* atomic masses. To understand relative masses, think about the setup in **Figure 4.** Eight pennies have the same mass as five nickels do. Thus, you could say that a single penny has a relative mass of 0.625 "nickel masses." Just as you can find the mass of a penny compared with the mass of a nickel, scientists have determined the masses of the elements relative to each other. Remember that atomic mass is given in units of *amu.* This means that it reflects an atom's mass relative to the mass of a carbon-12 atom. So, now you may ask why carbon's atomic mass on the periodic table is not exactly 12.

average atomic mass

the weighted average of the masses of all naturally occurring isotopes of an element

Most Elements Are Mixtures of Isotopes

You remember that *isotopes* are atoms that have different numbers of neutrons than other atoms of the same element do. So, isotopes have different atomic masses. The periodic table reports **average atomic mass,** a weighted average of the atomic mass of an element's isotopes. A *weighted average* takes into account the relative importance of each number in the average. Thus, if there is more of one isotope in a typical sample, it affects the average atomic mass more than an isotope that is less abundant does.

Figure 4
You can determine the mass of a penny relative to the mass of a nickel; eight pennies have the same mass as five nickels.

For example, carbon has two stable isotopes found in nature, carbon-12 and carbon-13. The average atomic mass of carbon takes into account the masses of both isotopes and their relative abundance. So, while the atomic mass of a carbon-12 atom is exactly 12 amu, any carbon sample will include enough carbon-13 atoms that the average mass of a carbon atom is 12.0107 amu.

Like carbon, most elements are a mixture of isotopes. In most cases, the fraction of each isotope is the same no matter where the sample comes from. Most average atomic masses can be determined to several decimal places. However, some elements have different percentages of isotopes depending on the source of the sample. This is true of *native* lead, or lead that occurs naturally on Earth. The average atomic mass of lead is given to only one decimal place because its composition varies so much from one sample to another.

If you know the abundance of each isotope, you can calculate the average atomic mass of an element. For example, the average atomic mass of native copper is a weighted average of the atomic masses of two isotopes, shown in **Figure 5.** The following sample problem shows how this calculation is made from data for the abundance of each of native copper's isotopes.

Figure 5
Native copper is a mixture of two isotopes. Copper-63 contributes 69.17% of the atoms, and copper-65 the remaining 30.83%.

SAMPLE PROBLEM E

Calculating Average Atomic Mass

The mass of a Cu-63 atom is 62.94 amu, and that of a Cu-65 atom is 64.93 amu. Using the data in **Figure 5,** find the average atomic mass of Cu.

1 Gather information.

- atomic mass of a Cu-63 atom = 62.94 amu
- abundance of Cu-63 = 69.17%
- atomic mass of Cu-65 = 64.93 amu
- abundance of Cu-65 = 30.83%
- average atomic mass of Cu = ? g

2 Plan your work.

The average atomic mass of an element is the sum of the contributions of the masses of each isotope to the total mass. This type of average is called a *weighted average*. The contribution of each isotope is equal to its atomic mass multiplied by the fraction of that isotope. (To change a percentage into a fraction, divide it by 100.)

Isotope	Percentage	Decimal fraction	Contribution
Copper-63	69.17%	0.6917	62.94×0.6917
Copper-65	30.83%	0.3083	64.93×0.3083

3 Calculate.

Average atomic mass is the sum of the individual contributions:

$$(62.94 \text{ amu} \times 0.6917) + (64.93 \text{ amu} \times 0.3083) = 63.55 \text{ amu}$$

4 Verify your results.

- The answer lies between 63 and 65, and the result is closer to 63 than it is to 65. This is expected because the isotope 63 makes a larger contribution to the average.
- Compare your answer with the value in the periodic table.

Practice problems on next page

PRACTICE HINT

In calculating average atomic masses, remember that the resulting value must be greater than the lightest isotope and less than the heaviest isotope.

PRACTICE

PROBLEM SOLVING SKILL

1. Calculate the average atomic mass for gallium if 60.00% of its atoms have a mass of 68.926 amu and 40.00% have a mass of 70.925 amu.
2. Calculate the average atomic mass of oxygen. Its composition is 99.76% of atoms with a mass of 15.99 amu, 0.038% with a mass of 17.00 amu, and 0.20% with a mass 18.00 amu.

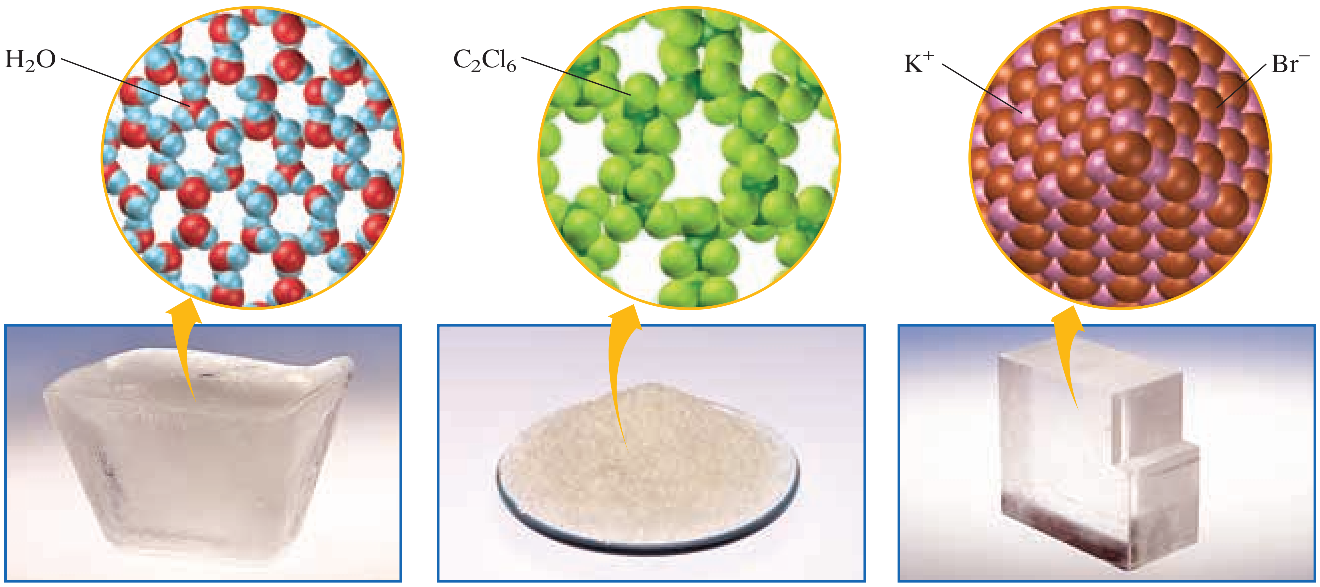
Chemical Formulas and Moles

Until now, when you needed to perform molar conversions, you were given the molar mass of compounds in a sample. Where does this molar mass of compounds come from? You can determine the molar mass of compounds the same way that you find the molar mass of individual elements—by using the periodic table.

Formulas Express Composition

The first step to finding a compound's molar mass is understanding what a chemical formula tells you. It tells you which elements, as well as how much of each, are present in a compound. The formula KBr shows that the compound is made up of potassium and bromide ions in a 1:1 ratio. The formula H_2O shows that water is made up of hydrogen and oxygen atoms in a 2:1 ratio. These ratios are shown in **Figure 6.**

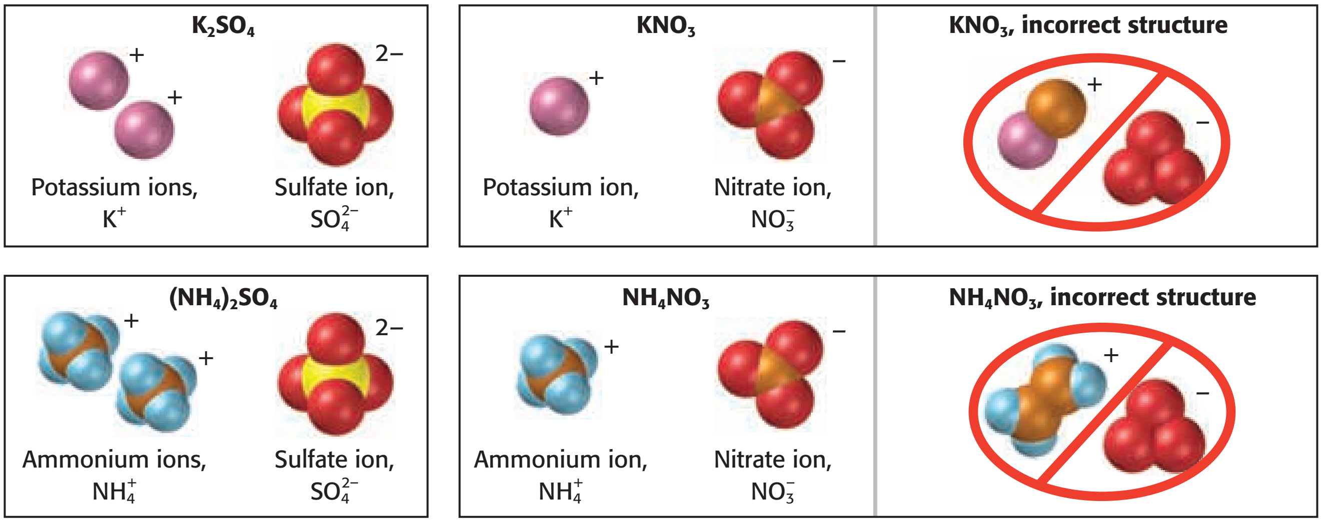
You have learned that covalent compounds, such as water and hexachloroethane, consist of molecules as units. Formulas for covalent compounds show both the elements and the number of atoms of each element in a molecule. Hexachloroethane has the formula C_2Cl_6. Each molecule has 8 atoms covalently bonded to each other. Ionic compounds aren't found as molecules, so their formulas do not show numbers of atoms. Instead, the formula shows the simplest ratio of cations and anions.



Water, H_2O | Hexachloroethane, C_2Cl_6 | Potassium bromide, KBr

Figure 6
Although any sample of a compound has many atoms and ions, the chemical formula gives a ratio of those atoms or ions.

Figure 7
The formula for a polyatomic ionic compound is the simplest ratio of cations to anions.



a Elements in polyatomic ions are bound together in a group and carry a characteristic charge.

b The formula for a compound with polyatomic ions shows how the atoms in each ion are bonded together.

c You cannot move atoms from one polyatomic ion to the next.

Formulas Give Ratios of Polyatomic Ions

The meaning of a formula does not change when polyatomic ions are involved. Potassium nitrate has the formula KNO_3. Just as the formula KBr indicates a 1:1 ratio of K^+ cations to Br^- anions, the formula KNO_3 indicates a ratio of one K^+ cation to one NO_3^- anion.

When a compound has polyatomic ions, such as those in **Figure 7,** look for the cations and anions. Formulas can tell you which elements make up polyatomic ions. For example, in the formula KNO_3, NO_3 is a nitrate ion, NO_3^-. KNO_3 does not have a KN^+ and an O_3^- ion. Similarly, the formula of ammonium nitrate is written NH_4NO_3, because NH_4 in a formula stands for the ammonium ion, NH_4^+, and NO_3 stands for a nitrate ion, NO_3^-. If it were written as $H_4N_2O_3$, the number of atoms would be correct. However, the formula would no longer clearly show which ions were in the substance and how many there were. The formula NH_4NO_3 shows that ammonium nitrate is made up of ammonium and nitrate ions in a 1:1 ratio.

Formulas Are Used to Calculate Molar Masses

A formula tells you what atoms (or ions) are present in an element or compound. So, from a formula you can find the mass of a mole of the substance, or its molar mass. The simplest formula for most elements is simply that element's symbol. For example, the symbol for silver is Ag. The molar mass of elements whose formulas are this simple equals the atomic mass of the element expressed in g/mol. So, the molar mass of silver is 107.87 g/mol. Diatomic elements have twice the number of atoms in each molecule, so their molecules have molar masses that are twice the molar mass of each atom. For example, the molar mass of Br_2 molecules is two times the molar mass of Br atoms (2×79.90 g/mol $= 159.80$ g/mol).

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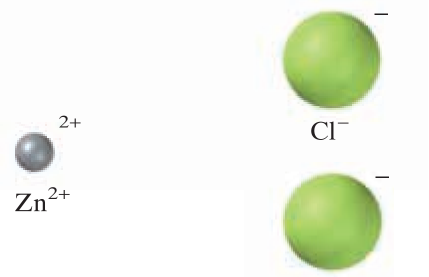
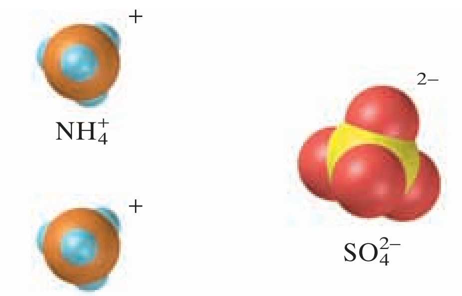
Let's say you want to determine the molar mass of a molecular compound. You must use the periodic table to find the molar mass of more than one element. The molar mass of a molecular compound is the sum of the masses of all the atoms in it expressed in g/mol. For example, one mole of H_2O molecules will have two moles of H and one mole of O. Thus, the compound's molar mass is equal to two times the molar mass of a H atom plus the molar mass of an O atom, or 18.02 g (2×1.01 g + 16.00 g).

Scientists also use the simplest formula to represent one mole of an ionic compound. They often use the term *formula unit* when referring to ionic compounds, because they are not found as single molecules. A formula unit of an ionic compound represents the simplest ratio of cations to anions. A formula unit of KBr is made up of one K^+ ion and one Br^- ion. One mole of an ionic compound has 6.022×10^{23} of these formula units. As with molecular compounds, the molar mass of an ionic compound is the sum of the masses of all the atoms in the formula expressed in g/mol. **Table 2** compares the formula units and molar masses of three ionic compounds. **Sample Problem F** shows how to calculate the molar mass of barium nitrate.

Table 2 Calculating Molar Mass for Ionic Compounds

Formula	Formula unit	Calculation of molar mass
$ZnCl_2$	 Zn^{2+}, Cl^-, Cl^-	1 Zn = 1×65.39 g/mol = 65.39 g/mol + 2 Cl = 2×35.45 g/mol = 70.90 g/mol $ZnCl_2$ = 136.29 g/mol
$ZnSO_4$	 Zn^{2+}, SO_4^{2-}	1 Zn = 1×65.39 g/mol = 65.39 g/mol 1 S = 1×32.07 g/mol = 32.07 g/mol + 4 O = 4×16.00 g/mol = 64.00 g/mol $ZnSO_4$ = 161.46 g/mol
$(NH_4)_2SO_4$	 NH_4^+, NH_4^+, SO_4^{2-}	2 N = 2×14.01 g/mol = 28.02 g/mol 8 H = 8×1.01 g/mol = 8.08 g/mol 1 S = 1×32.07 g/mol = 32.07 g/mol + 4 O = 4×16.00 g/mol = 64.00 g/mol $(NH_4)_2SO_4$ = 132.17 g/mol

SAMPLE PROBLEM F

Calculating Molar Mass of Compounds

Find the molar mass of barium nitrate.

1 Gather information.

- simplest formula of ionic barium nitrate: $Ba(NO_3)_2$
- molar mass of $Ba(NO_3)_2$ = ? g/mol

2 Plan your work.

- Find the number of moles of each element in 1 mol $Ba(NO_3)_2$. Each mole has:

  1 mol Ba
  2 mol N
  6 mol O

- Use the periodic table to find the molar mass of each element in the formula.

$$\text{molar mass of Ba} = 137.33 \text{ g/mol}$$
$$\text{molar mass of N} = 14.01 \text{ g/mol}$$
$$\text{molar mass of O} = 16.00 \text{ g/mol}$$

3 Calculate.

- Multiply the molar mass of each element by the number of moles of each element. Add these masses to get the total molar mass of $Ba(NO_3)_2$.

$$\begin{aligned}
\text{mass of 1 mol Ba} &= 1 \times 137.33 \text{ g/mol} = 137.33 \text{ g/mol} \\
\text{mass of 2 mol N} &= 2 \times 14.01 \text{ g/mol} = 28.02 \text{ g/mol} \\
+\ \text{mass of 6 mol O} &= 6 \times 16.00 \text{ g/mol} = 96.00 \text{ g/mol} \\
\hline
\text{molar mass of } Ba(NO_3)_2 &= 261.35 \text{ g/mol}
\end{aligned}$$

4 Verify your result.

- The answer has the correct units. The sum of the molar masses of elements can be approximated as $140 + 30 + 100 = 270$, which is close to the calculated value.

PRACTICE HINT

Use the same methods for molecular compounds, but use the molecular formula in place of a formula unit.

PRACTICE

1 Find the molar mass for each of the following compounds:

a. CsI **c.** $C_{12}H_{22}O_{11}$ **e.** $HC_2H_3O_2$

b. $CaHPO_4$ **d.** I_2 **f.** $Mg_3(PO_4)_2$

2 Write the formula and then find the molar mass.

a. sodium hydrogen carbonate **e.** iron(III) hydroxide

b. cerium hexaboride **f.** tin(II) chloride

c. magnesium perchlorate **g.** tetraphosphorus decoxide

d. aluminum sulfate **h.** iodine monochloride

continued on next page

PROBLEM SOLVING SKILL

PRACTICE

PROBLEM SOLVING SKILL

3 **a.** Find the molar mass of toluene, $C_6H_5CH_3$.

b. Find the number of moles in 7.51 g of toluene.

4 **a.** Find the molar mass of cisplatin, $PtCl_2(NH_3)_2$, a cancer therapy chemical.

b. Find the mass of 4.115×10^{21} formula units of cisplatin.

2 Section Review

UNDERSTANDING KEY IDEAS

1. What is a weighted average?
2. On the periodic table, the average atomic mass of carbon is 12.01 g. Why is it not exactly 12.00?
3. What is the simplest formula for cesium carbonate?
4. What ions are present in cesium carbonate?
5. What is the ratio of N and H atoms in NH_3?
6. What is the ratio of calcium and chloride ions in $CaCl_2$?
7. Why is the simplest formula used to determine the molar mass for ionic compounds?

PRACTICE PROBLEMS

8. Calculate the average atomic mass of chromium. Its composition is: 83.79% with a mass of 51.94 amu; 9.50% with a mass of 52.94 amu; 4.35% with a mass of 49.95 amu; 2.36% with a mass of 53.94 amu.
9. Element X has two isotopes. One has a mass of 10.0 amu and an abundance of 20.0%. The other has a mass of 11.0 amu and an abundance of 80.0%. Estimate the average atomic mass. What element is it?
10. Find the molar mass.
    - **a.** CsCl
    - **b.** $KClO_3$
    - **c.** $C_6H_{12}O_6$
    - **d.** $(NH_4)_2HPO_4$
    - **e.** $C_2H_5NO_2$
11. Determine the formula, the molar mass, and the number of moles in 2.11 g of each of the following compounds.
    - **a.** strontium sulfide
    - **b.** phosphorus trifluoride
    - **c.** zinc acetate
    - **d.** mercury(II) bromate
    - **e.** calcium nitrate
12. Find the molar mass and the mass of 5.0000 mol of each of the following compounds.
    - **a.** calcium acetate, $Ca(C_2H_3O_2)_2$
    - **b.** iron(II) phosphate, $Fe_3(PO_4)_2$
    - **c.** saccharin, $C_7H_5NO_3S$, a sweetener
    - **d.** acetylsalicylic acid, $C_9H_8O_4$, or aspirin

CRITICAL THINKING

13. In the periodic table, the atomic mass of fluorine is given to 9 significant figures, whereas oxygen is given to only 6. Why? (Hint: fluorine has only one isotope.)
14. **Figure 6** shows many K^+ and Br^- ions. Why is the formula not written as $K_{20}Br_{20}$?
15. Why don't scientists use HO as the formula for hydrogen peroxide, H_2O_2?
16. **a.** How many atoms of H are in a formula unit of $(NH_4)_2SO_4$?

    b. How many atoms of H are in 1 mol of $(NH_4)_2SO_4$?

SECTION

3 Formulas and Percentage Composition

KEY TERMS

- **percentage composition**
- **empirical formula**
- **molecular formula**

OBJECTIVES

1. **Determine** a compound's empirical formula from its percentage composition.
2. **Determine** the molecular formula or formula unit of a compound from its empirical formula and its formula mass.
3. **Calculate** percentage composition of a compound from its molecular formula or formula unit.

Using Analytical Data

Scientists synthesize new compounds for many uses. Once they make a new product, they must check its identity. One way is to carry out a chemical analysis that provides a **percentage composition.** For example, in 1962, two chemists made a new compound from xenon and fluorine. Before 1962, scientists thought that xenon did not form compounds. The scientists analyzed their surprising find. They found that it had a percentage composition of 63.3% Xe and 36.7% F, which is the same as that for the formula XeF_4. Percentage composition not only helps verify a substance's identity but also can be used to compare the ratio of masses contributed by the elements in two substances, as in **Figure 8.**

percentage composition
the percentage by mass of each element in a compound

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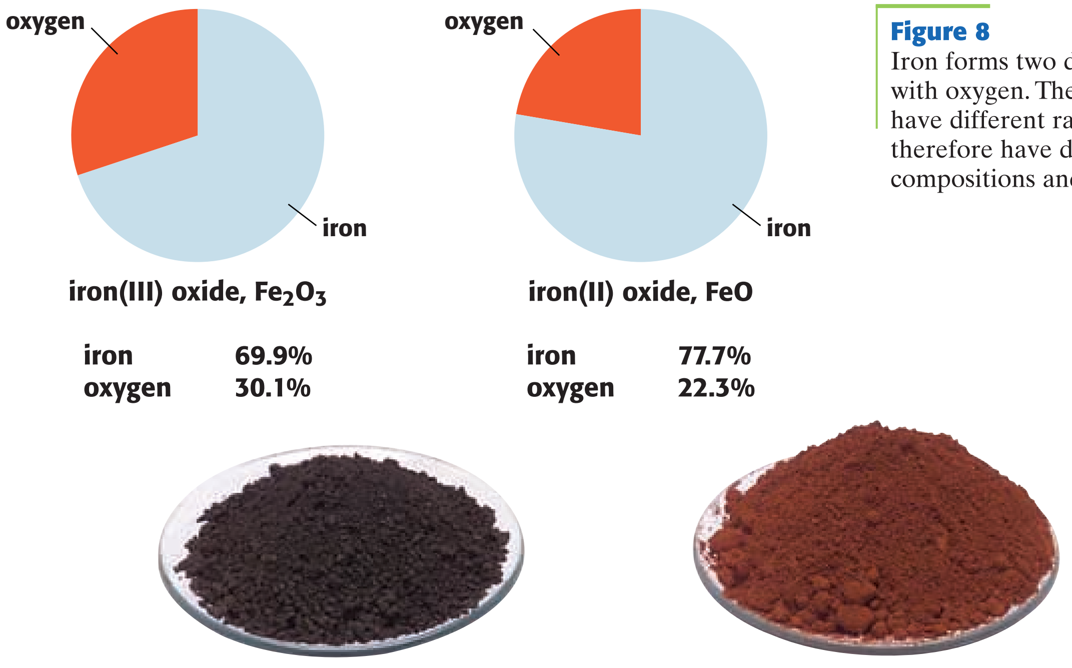


Figure 8
Iron forms two different compounds with oxygen. The two compounds have different ratios of atoms and therefore have different percentage compositions and different properties.

empirical formula

a chemical formula that shows the composition of a compound in terms of the relative numbers and kinds of atoms in the simplest ratio

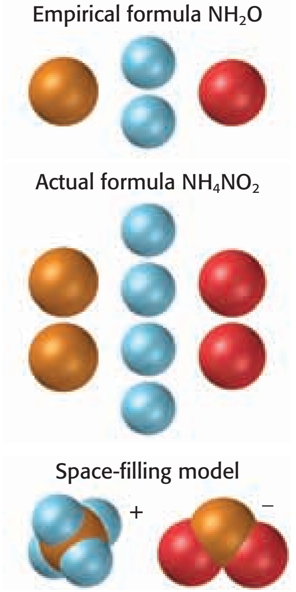


Figure 9
The empirical formula for ammonium nitrite is NH_2O. Its actual formula has 1 ammonium ion, NH_4^+, and 1 nitrite ion, NO_2^-.

Determining Empirical Formulas

Data for percentage composition allow you to calculate the simplest ratio among the atoms found in a compound. The **empirical formula** shows this simplest ratio. For example, ammonium nitrite, shown in **Figure 9,** has the actual formula NH_4NO_2 and is made up of ammonium ions, NH_4^+, and nitrite ions, NO_2^-, in a 1:1 ratio.

But if a chemist does an elemental analysis, she will find the empirical formula to be NH_2O, because it shows the *simplest ratio* of the elements. For some other compounds, the empirical formula and the actual formula are the same.

Let's say that you want to find an empirical formula from the percentage composition. First, convert the mass percentage of each element to grams. Second, convert from grams to moles using the molar mass of each element as a conversion factor. (Keep in mind that a formula for a compound can be read as a number of atoms or as a number of moles.) Third, as shown in **Sample Problem G,** compare these amounts in moles to find the simplest whole-number ratio among the elements in the compound.

To find this ratio, divide each amount by the smallest of all the amounts. This process will give a subscript of 1 for the atoms present in the smallest amount. Finally, you may need to multiply by a number to convert all subscripts to the smallest whole numbers. The final numbers you get are the subscripts in the empirical formula. For example, suppose the subscripts were 1.33, 2, and 1. Multiplication by 3 gives subscripts of 4, 6, and 3.

SAMPLE PROBLEM G

Determining an Empirical Formula from Percentage Composition

Chemical analysis of a liquid shows that it is 60.0% C, 13.4% H, and 26.6% O by mass. Calculate the empirical formula of this substance.

1 Gather information.

- percentage C = 60.0%
- percentage H = 13.4%
- percentage O = 26.6%
- empirical formula = $C_?H_?O_?$

2 Plan your work.

- Assume that you have a 100.0 g sample of the liquid, and convert the percentages to grams.

$$\text{for C: } 60.0\% \times 100.0 \text{ g} = 60.0 \text{ g C}$$

$$\text{for H: } 13.4\% \times 100.0 \text{ g} = 13.4 \text{ g H}$$

$$\text{for O: } 26.6\% \times 100.0 \text{ g} = 26.6 \text{ g O}$$

- To convert the mass of each element into the amount in moles, you must multiply by the proper conversion factor, which is the reciprocal of the molar mass. Find molar mass by using the periodic table.

molar mass of C: 12.01 g/mol

molar mass of H: 1.01 g/mol

molar mass of O: 16.00 g/mol

3 Calculate.

- Calculate the amount in moles of C, H, and O. Round the answers to the correct number of significant figures.

$$60.0\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 5.00\ \text{mol C}$$

$$13.4\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}} = 13.3\ \text{mol H}$$

$$26.6\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 1.66\ \text{mol O}$$

- At this point the formula can be written as $C_5H_{13.3}O_{1.66}$, but you know that subscripts in chemical formulas are usually whole numbers.
- To begin the conversion to whole numbers, divide all subscripts by the smallest subscript, 1.66. This will make at least one of the subscripts a whole number, 1.

$$\frac{5.00\ \text{mol C}}{1.66} = 3.01\ \text{mol C}$$

$$\frac{13.3\ \text{mol H}}{1.66} = 8.01\ \text{mol H}$$

$$\frac{1.66\ \text{mol O}}{1.66} = 1.00\ \text{mol O}$$

- These numbers can be assumed to be the whole numbers 3, 8 and 1. The empirical formula is therefore C_3H_8O.

4 Verify your result.

Verify your answer by calculating the percentage composition of C_3H_8O. If the result agrees with the composition stated in the problem, then the formula is correct.

PRACTICE HINT

When you get fractions for the first calculation of subscripts, think about how you can turn these into whole numbers. For example:

- the subscript 1.33 is roughly $1\frac{1}{3}$, so it will give the whole number 4 when multiplied by 3
- the subscript 0.249 is roughly $\frac{1}{4}$, so it will give the whole number 1 when multiplied by 4
- the subscript 0.74 is roughly $\frac{3}{4}$, so it will give the whole number 3 when multiplied by 4

PRACTICE

Determine the empirical formula for each substance.

1. A dead alkaline battery is found to contain a compound of Mn and O. Its analysis gives 69.6% Mn and 30.4% O.
2. A compound is 38.77% Cl and 61.23% O.
3. Magnetic iron oxide is 72.4% iron and 27.6% oxygen.
4. A liquid compound is 18.0% C, 2.26% H, and 79.7% Cl.

PROBLEM SOLVING SKILL

molecular formula

a chemical formula that shows the number and kinds of atoms in a molecule, but not the arrangement of the atoms

Molecular Formulas Are Multiples of Empirical Formulas

The formula for an ionic compound shows the simplest whole-number ratio of the large numbers of ions in a crystal of the compound. The formula $Ca_3(PO_4)_2$ shows that the ratio of Ca^{2+} ions to PO_4^{3-} ions is 3:2.

Molecular compounds, on the other hand, are made of single molecules. Some molecular compounds have the same molecular and empirical formulas. Examples are water, H_2O, and nitric acid, HNO_3. But for many molecular compounds the **molecular formula** is a whole-number multiple of the empirical formula. Both kinds of formulas are just two different ways of representing the composition of the same molecule.

The molar mass of a compound is equal to the molar mass of the empirical formula times a whole number, n. There are several experimental techniques for finding the molar mass of a molecular compound even though the compound's chemical composition and formula are unknown. If you divide the experimental molar mass by the molar mass of the empirical formula, you can figure out the value of n needed to scale the empirical formula up to give the molecular formula.

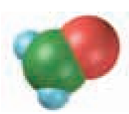
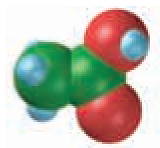
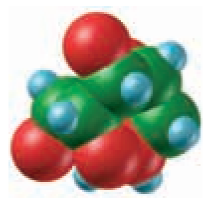
Figure 10
The formula for glucose, which is found in many sports drinks, is $C_6H_{12}O_6$.

Think about the three compounds in **Table 3**—formaldehyde, acetic acid, and glucose, which is shown in **Figure 10.** Each has the empirical formula CH_2O. However, acetic acid has a molecular formula that is twice the empirical formula. The molecular formula for glucose is six times the empirical formula. The relationship is shown in the following equation.

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

In general, the molecular formula is a whole-number multiple of the empirical formula. For formaldehyde, $n = 1$, for acetic acid, $n = 2$, and for glucose, $n = 6$. In some cases, n may be a very large number.

Table 3 Comparing Empirical and Molecular Formulas

Compound	Empirical formula	Molecular formula	Molar mass (g)	Space-filling model
Formaldehyde	CH_2O	CH_2O • same as empirical formula • $n = 1$	30.03	
Acetic acid	CH_2O	$C_2H_4O_2$ ($HC_2H_3O_2$) • 2 × empirical formula • $n = 2$	60.06	
Glucose	CH_2O	$C_6H_{12}O_6$ • 6 × empirical formula • $n = 6$	180.18	

SAMPLE PROBLEM H

Determining a Molecular Formula from an Empirical Formula

The empirical formula for a compound is P_2O_5. Its experimental molar mass is 284 g/mol. Determine the molecular formula of the compound.

1 Gather information.

- empirical formula = P_2O_5
- molar mass of compound = 284 g/mol
- molecular formula = ?

2 Plan your work.

- Find the molar mass of the empirical formula using the molar masses of the elements from the periodic table.

$$\text{molar mass of P} = 30.97 \text{ g/mol}$$
$$\text{molar mass of O} = 16.00 \text{ g/mol}$$

3 Calculate.

- Find the molar mass of the empirical formula, P_2O_5.

$$\begin{array}{rr} 2 \times \text{molar mass of P} = & 61.94 \text{ g/mol} \\ + \; 5 \times \text{molar mass of O} = & 80.00 \text{ g/mol} \\ \hline \text{molar mass of } P_2O_5 = & 141.94 \text{ g/mol} \end{array}$$

- Solve for n, the factor multiplying the empirical formula to get the molecular formula.

$$n = \frac{\text{experimental molar mass of compound}}{\text{molar mass of empirical formula}}$$

- Substitute the molar masses into this equation, and solve for n.

$$n = \frac{284 \text{ g/mol}}{141.94 \text{ g/mol}} = 2.00 = 2$$

- Multiply the empirical formula by this factor to get the answer.

$$n(\text{empirical formula}) = 2(P_2O_5) = P_4O_{10}$$

4 Verify your result.

- The molar mass of P_4O_{10} is 283.88 g/mol. It is equal to the experimental molar mass.

PRACTICE HINT

In some cases, you can figure out the factor n by just looking at the numbers. For example, let's say you noticed that the experimental molar mass was almost exactly twice as much as the molar mass of the empirical formula (as in this problem). That means n must be 2.

PRACTICE

1. A compound has an experimental molar mass of 78 g/mol. Its empirical formula is CH. What is its molecular formula?
2. A compound has the empirical formula CH_2O. Its experimental molar mass is 90.0 g/mol. What is its molecular formula?
3. A brown gas has the empirical formula NO_2. Its experimental molar mass is 46 g/mol. What is its molecular formula?

PROBLEM SOLVING SKILL

Chemical Formulas Can Give Percentage Composition

If you know the chemical formula of any compound, then you can calculate the percentage composition. From the subscripts, you can determine the mass contributed by each element and add these to get the molar mass. Then, divide the mass of each element by the molar mass. Multiply by 100 to find the percentage composition of that element.

Think about the two compounds shown in **Figure 11.** Carbon dioxide, CO_2, is a harmless gas that you exhale, while carbon monoxide, CO, is a poisonous gas present in car exhaust. The percentage composition of carbon dioxide, CO_2, is calculated as follows.

$$\begin{array}{r} 1 \text{ mol} \times 12.01 \text{ g C/mol} = 12.01 \text{ g C} \\ + 2 \text{ mol} \times 16.00 \text{ g O/mol} = 32.00 \text{ g O} \\ \hline \text{mass of 1 mol } CO_2 = 44.01 \text{ g} \end{array}$$

$$\% \text{ C in } CO_2 = \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} \times 100 = 27.29\%$$

$$\% \text{ O in } CO_2 = \frac{32.00 \text{ g O}}{44.01 \text{ g } CO_2} \times 100 = 72.71\ \%$$

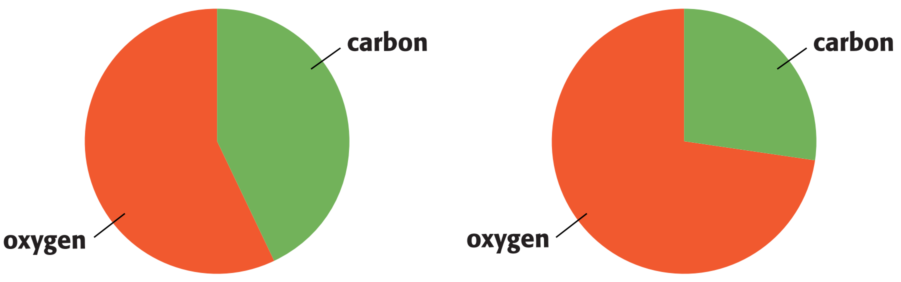
The percentage composition of carbon monoxide, CO, is calculated as follows.

$$\begin{array}{r} 1 \text{ mol} \times 12.01 \text{ g C/mol} = 12.01 \text{ g C} \\ + 1 \text{ mol} \times 16.00 \text{ g O/mol} = 16.00 \text{ g O} \\ \hline \text{mass of 1 mol CO} = 28.01 \text{ g} \end{array}$$

$$\% \text{ C in CO} = \frac{12.01 \text{ g C}}{28.01 \text{ g CO}} \times 100 = 42.88\%$$

$$\% \text{ O in CO} = \frac{16.00 \text{ g O}}{28.01 \text{ g CO}} \times 100 = 57.71\ \%$$

Figure 11
Carbon monoxide and carbon dioxide are both made up of the same elements, but they have different percentage compositions.



carbon monoxide, CO

carbon	**42.88%**
oxygen	**57.12%**

carbon dioxide, CO_2

carbon	**27.29%**
oxygen	**72.71%**

SAMPLE PROBLEM I

Using a Chemical Formula to Determine Percentage Composition

Calculate the percentage composition of copper(I) sulfide, a copper ore called chalcocite.

1 Gather information.

- name and formula of the compound: copper(I) sulfide, Cu_2S
- percentage composition: $\%Cu = ?$, $\%S = ?$

2 Plan your work.

To determine the molar mass of copper(I) sulfide, find the molar mass of the elements copper and sulfur using the periodic table.

$$\text{molar mass of Cu} = 63.55 \text{ g/mol}$$
$$\text{molar mass of S} = 32.07 \text{ g/mol}$$

3 Calculate.

- Find the masses of 2 mol Cu and 1 mol S. Use these masses to find the molar mass of Cu_2S.

$$\begin{array}{r} 2 \text{ mol} \times 63.55 \text{ g Cu/mol} = 127.10 \text{ g Cu} \\ + \; 1 \text{ mol} \times 32.07 \text{ g S/mol} = 32.07 \text{ g S} \\ \hline \text{molar mass of } Cu_2S = 159.17 \text{ g/mol} \end{array}$$

- Calculate the fraction that each element contributes to the total mass. Do this by dividing the total mass contributed by that element by the total mass of the compound. Convert the fraction to a percentage by multiplying by 100.

$$\text{mass \% Cu} = \frac{\text{mass of 2 mol Cu}}{\text{molar mass of } Cu_2S} \times 100$$

$$\text{mass \% S} = \frac{\text{mass of 1 mol S}}{\text{molar mass of } Cu_2S} \times 100$$

- Substitute the masses into the equations above. Round the answers you get on the calculator to the correct number of significant figures.

$$\text{mass \% Cu} = \frac{127.10 \text{ g Cu}}{159.17 \text{ g } Cu_2S} \times 100 = 79.852\% \text{ Cu}$$

$$\text{mass \% S} = \frac{32.07 \text{ g S}}{159.17 \text{ g } Cu_2S} \times 100 = 20.15\% \text{ S}$$

4 Verify your result.

- Add the percentages. The sum should be near 100%.

$$79.852\% + 20.15\% = 100.00\%$$

Practice problems on next page

PRACTICE HINT

Sometimes, rounding gives a sum that differs slightly from 100%. This is expected. (However, if you find a sum that differs significantly, such as 112%, you have made an error.)

PROBLEM SOLVING SKILL

PRACTICE

1. Calculate the percentage composition of Fe_3C, a compound in cast iron.
2. Calculate the percentage of both elements in sulfur dioxide.
3. Calculate the percentage composition of ammonium nitrate, NH_4NO_3.
4. Calculate the percentage composition of each of the following:
   a. $SrBr_2$
   b. $CaSO_4$
   c. $Mg(CN)_2$
   d. $Pb(CH_3COO)_2$
5. **a.** Calculate the percentage of each element in acetic acid, $HC_2H_3O_2$, and glucose, $C_6H_{12}O_6$.
   b. These two substances have the same empirical formula. What would you expect the percentage composition of the empirical formula to be?

3 Section Review

UNDERSTANDING KEY IDEAS

1. a. Suppose you know that a compound is 11.2% H and 88.8% O. What information do you need to determine the empirical formula?
   b. What additional information do you need to determine the molecular formula?

2. Isooctane has the molecular formula C_8H_{18}. What is its empirical formula?

3. What information do you need to calculate the percentage composition of CF_4?

PRACTICE PROBLEMS

4. Determine the empirical formula.
   a. The analysis of a compound shows that it is 9.2% B and 90.8% Cl.
   b. An analysis shows that a compound is 50.1% S and 49.9% O.
   c. The analysis of a compound shows that it is 27.0% Na, 16.5% N, and 56.5% O.

5. The experimental molar mass of the compound in item 4b is 64 g/mol. What is the compound's molecular formula?

6. Determine the formula, and then calculate the percentage composition.
   a. calcium sulfate
   b. silicon dioxide
   c. silver nitrate
   d. nitrogen monoxide

7. Calculate the percentage composition.
   a. silver acetate, $AgC_2H_3O_2$
   b. lead(II) chlorate, $Pb(ClO_3)_2$
   c. iron(III) sulfate, $Fe_2(SO_4)_3$
   d. copper(II) sulfate, $CuSO_4$

CRITICAL THINKING

8. When you determine the empirical formula of a compound from analytical data, you seldom get exact whole numbers for the subscripts. Explain why.

9. An amino acid has the molecular formula $C_2H_5NO_2$. What is the empirical formula?

10. A compound has the empirical formula CH_2O. Its experimental molar mass is 45 g/mol. Is it possible to calculate the molecular formula with the information given?

82
Pb
Lead
207.2
$[Xe]4f^{14}5d^{10}6s^{2}6p^{2}$

Where Is Pb?

Earth's crust:
< 0.01% by mass

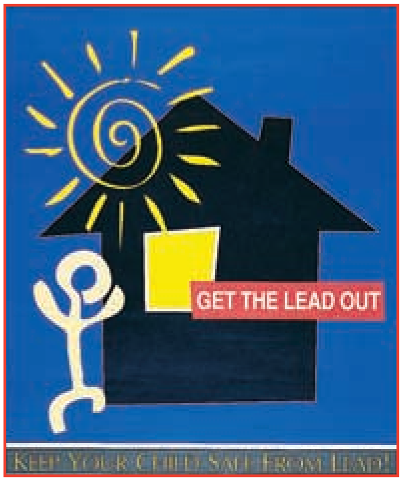
Element Spotlight

Get the Lead Out

Humans have known for many centuries that lead is toxic, but it is still used in many common materials. High levels of lead were used in white paints until the 1940s. Since then, the lead compounds in paints have gradually been replaced with less toxic titanium dioxide. However, many older buildings still have significant amounts of lead paint, and many also have lead solder in their water pipes.

Lead poisoning is caused by the absorption of lead through the digestive tract, lungs, or skin. Children living in older homes are especially susceptible to lead poisoning. Children eat paint chips that contain lead because the paint has a sweet taste.

The hazards of lead poisoning can be greatly reduced by introducing programs that increase public awareness, removing lead-based paint from old buildings, and screening children for lead exposure.



Posters such as this one are part of public-awareness programs to reduce the hazards of lead poisoning.

Industrial Uses

- The largest industrial use of lead is in the manufacture of storage batteries.
- Solder used for joining metals is often an alloy of lead and tin.
- Other lead alloys are used to make bearings for gasoline and diesel engines, type metal for printing, corrosion-resistant cable coverings, and ammunition.
- Lead sheets and lead bricks are used to shield workers and sensitive objects from X rays.

Real-World Connection Lead inside the human body interferes with the production of red blood cells and can cause damage to the kidneys, liver, brain, and other organs.

A Brief History

3000 BCE — 1000 BCE — 1 CE — 1000 CE

3000 BCE: Egyptians refine and use lead to make art figurines.

600 BCE: Lead ore deposits are discovered near Athens; they are mined until the second century CE.

60 BCE: Romans begin making lead pipes, lead sheets for waterproofing roofs, and lead crystal.

1977: The U.S. government restricts lead content in paint.

Questions

1. How do you perform tests for lead in paint, soil, and water? Present a report that explains how the tests work.

2. Research the laws regarding the recycling of storage batteries that contain lead.

internet connect
www.scilinks.org
Topic : Lead
SciLinks code: HW4074
SciLinks Maintained by the National Science Teachers Association

7 CHAPTER HIGHLIGHTS

KEY IDEAS

SECTION ONE Avogadro's Number and Molar Conversions

- Avogadro's number, 6.022×10^{23} units/mol, is the number of units (atoms, ions, molecules, formula units, etc.) in 1 mol of any substance.
- Avogadro's number is used to convert from number of moles to number of particles or vice versa.
- Conversions between moles and mass require the use of molar mass.
- The molar mass of a monatomic element is the number of grams numerically equal to the atomic mass on the periodic table.

SECTION TWO Relative Atomic Mass and Chemical Formulas

- The average atomic mass of an element is the average mass of the element's isotopes, weighted by the percentage of their natural abundance.
- Chemical formulas reveal composition. The subscripts in the formula give the number of atoms of a given element in a molecule or formula unit of a compound or diatomic element.
- Formulas are used to calculate molar masses of compounds.

SECTION THREE Formulas and Percentage Composition

- Percentage composition gives the relative contribution of each element to the total mass of one molecule or formula unit.
- An empirical formula shows the elements and the smallest whole-number ratio of atoms or ions that are present in a compound. It can be found by using the percentage composition.
- The molecular formula is determined from the empirical formula and the experimentally determined molar mass.
- Chemical formulas can be used to calculate percentage composition.

KEY TERMS

mole
Avogadro's number
molar mass

average atomic mass

percentage composition
empirical formula
molecular formula

KEY SKILLS

Working Practice Problems
Skills Toolkit 2 p. 227

Converting Between Amount in Moles and Number of Particles
Skills Toolkit 1 p. 226
Sample Problem A p. 228
Sample Problem B p. 229

Converting Between Mass, Amount, and Number of Particles
Skills Toolkit 3 p. 230
Sample Problem C p. 231
Sample Problem D p. 232

Calculating Average Atomic Mass
Sample Problem E p. 235

Calculating Molar Mass of Compounds
Sample Problem F p. 239

Determining an Empirical Formula from Percentage Composition
Sample Problem G p. 242

Determining a Molecular Formula from an Empirical Formula
Sample Problem H p. 245

Using a Chemical Formula to Determine Percentage Composition
Sample Problem I p. 247

CHAPTER REVIEW 7

USING KEY TERMS

1. Distinguish between Avogadro's number and the mole.
2. What term is used to describe the mass in grams of 1 mol of a substance?
3. Why is the ratio between the empirical formula and the molecular formula a whole number?
4. What do you need to calculate the percentage composition of a substance?
5. Explain the difference between atomic mass and average atomic mass.

UNDERSTANDING KEY IDEAS

Avogadro's Number and Molar Conversions

6. What particular isotope is the basis for defining the atomic mass unit and the mole?
7. How would you determine the number of molecules in 3 mol of oxygen, O_2?
8. What conversion factor do you use in converting number of moles into number of formula units?
9. How is molar mass of an element used to convert from number of moles to mass in grams?
10. What result do you get when you multiply the number of moles of a sample by the following conversion factor?

$$\frac{\text{g of element}}{\text{1 mol element}}$$

11. You convert 10 mol of a substance to grams. Is the number in the answer larger or smaller than 10 g?

Relative Atomic Mass and Chemical Formulas

12. Which has the greater number of molecules: 10 g of N_2 or 10 g of O_2?
13. How is average atomic mass determined from isotopic masses?
14. For an element, what is the relationship between atomic mass and molar mass?
15. How do you determine the molar mass of a compound?

Formulas and Percentage Composition

16. What information does percentage composition reveal about a compound?
17. Summarize briefly the process of using empirical formula and the value for experimental molar mass to determine the molecular formula.
18. When you calculate the percentage composition of a compound from both the empirical formula and the molecular formula, why are the two results identical?

PRACTICE PROBLEMS

PROBLEM SOLVING SKILL

Sample Problem A Converting Amount in Moles to Number of Particles

19. How many sodium ions in 2.00 mol of NaCl?
20. How many molecules in 2.00 mol of sucrose, $C_{12}H_{22}O_{11}$?

21. How many atoms are in the 1.25×10^{-2} mol of mercury within the bulb of a thermometer?

22. a. How many formula units are there in a 3.12 mol sample of $MgCl_2$?

b. How many Cl^- ions are there in the sample?

Sample Problem B Converting Number of Particles to Amount in Moles

23. How many moles of magnesium oxide are there in 2.50×10^{25} formula units of MgO?

24. A sample has 7.51×10^{24} molecules of benzene, C_6H_6. How many moles is this?

25. How many moles are in a sample having 9.3541×10^{13} particles?

26. How many moles of sodium ions are there in a sample of salt water that contains 4.11×10^{22} Na^+ ions?

27. How many moles are equal to 3.6×10^{23} molecules of oxygen gas, O_2?

Sample Problem C Converting Number of Particles to Mass

28. How many grams are present in 4.336×10^{24} formula units of table salt, NaCl, whose molar mass is 58.44 g/mol?

29. A scientist collects a sample that has 2.00×10^{14} molecules of carbon dioxide gas. How many grams is this, given that the molar mass of CO_2 is 44.01 g/mol?

30. What is the mass in grams of a sample of $Fe_2(SO_4)_3$ that contains 3.59×10^{23} sulfate ions, SO_4^{2-} ? The molar mass of $Fe_2(SO_4)_3$ is 399.91 g/mol.

31. Calculate the mass in grams of 2.55 mol of oxygen gas, O_2 (molar mass of O_2 = 32.00 g/mol).

32. How many grams are in 2.7 mol of table salt, NaCl (molar mass of NaCl = 58.44 g/mol)?

33. Calculate the mass of each of the following samples:

a. 0.500 mol I_2 (molar mass of I_2 = 253.80 g/mol)

b. 2.82 mol PbS (molar mass of PbS = 239.3 g/mol)

c. 4.00 mol of C_4H_{10} (molar mass of C_4H_{10} = 58.14 g/mol)

34. How many grams are in each of the following samples?

a. 1.000 mol NaCl (molar mass of NaCl = 58.44 g/mol)

b. 2.000 mol H_2O (molar mass of H_2O = 18.02 g/mol)

c. 3.5 mol $Ca(OH)_2$ (molar mass of $Ca(OH)_2$ = 74.10 g/mol)

Sample Problem D Converting Mass to Number of Particles

35. How many atoms of gold are there in a pure gold ring with a mass of 10.6 g?

36. How many formula units are there in 302.48 g of zinc chloride, $ZnCl_2$? The molar mass of zinc chloride is 136.29 g/mol.

37. Naphthalene, $C_{10}H_8$, an ingredient in mothballs, has a molar mass of 128.18 g/mol. How many molecules of naphthalene are in a mothball that has 2.000 g of naphthalene.

38. How many moles of compound are in each of the following samples:

a. 6.60 g $(NH_4)_2SO_4$ (molar mass of $(NH_4)_2SO_4$ = 132.17 g/mol)

b. 4.5 kg of $Ca(OH)_2$ (molar mass of $Ca(OH)_2$ = 74.10 g/mol)

39. Ibuprofen, $C_{13}H_{18}O_2$, an active ingredient in pain relievers has a molar mass of 206.31 g/mol. How many moles of ibuprofen are in a bottle that contains 33 g of ibuprofen?

40. How many moles of $NaNO_2$ are there in a beaker that contains 0.500 kg of $NaNO_2$ (molar mass of $NaNO_2$ = 69.00 g/mol)?

41. How many moles of propane are in a pressure container that has 2.55 kg of propane, C_3H_8 (molar mass of C_3H_8 = 44.11 g/mol)?

Sample Problem E Calculating Average Atomic Mass

42. Naturally occurring silver is composed of two isotopes: Ag-107 is 51.35% with a mass of 106.905092 amu, and the rest is Ag-109 with a mass of 108.9044757 amu. Calculate the average atomic mass of silver.

43. The element bromine is distributed between two isotopes. The first, amounting to 50.69%, has a mass of 78.918 amu. The second, amounting to 49.31%, has a mass of 80.916 amu. Calculate the average atomic mass of bromine.

44. Calculate the average atomic mass of iron. Its composition is 5.90% with a mass of 53.94 amu, 91.72% with a mass of 55.93 amu, 2.10% with a mass of 56.94 amu, and 0.280% with a mass of 57.93 amu.

Sample Problem F Calculating Molar Mass of Compounds

45. Find the molar mass of the following compounds:
    a. lithium chloride
    b. copper(I) cyanide
    c. potassium dichromate
    d. magnesium nitrate
    e. tetrasulfur tetranitride

46. What is the molar mass of the phosphate ion, PO_4^{3-}?

47. Find the molar mass of isopropyl alcohol, C_3H_7OH, used as rubbing alcohol.

48. What is the molar mass of the amino acid glycine, $C_2H_5NO_2$?

Sample Problem G Determining an Empirical Formula from Percentage Composition

49. A compound of silver has the following analytical composition: 63.50% Ag, 8.25% N, and 28.25% O. Calculate the empirical formula.

50. An oxide of phosphorus is 56.34% phosphorus, and the rest is oxygen. Calculate the empirical formula for this compound.

Sample Problem H Determining a Molecular Formula from an Empirical Formula

51. The empirical formula of the anticancer drug altretamine is $C_3H_6N_2$. The experimental molar mass is 210 g/mol. What is its molecular formula?

52. Benzene has the empirical formula CH and an experimental molar mass of 78 g/mol. What is its molecular formula?

53. Determine the molecular formula for a compound with the empirical formula CoC_4O_4 and a molar mass of 341.94 g/mol.

54. Oleic acid has the empirical formula $C_9H_{17}O$. If the experimental molar mass is 282 g/mol, what is the molecular formula of oleic acid?

Sample Problem I Using a Chemical Formula to Determine Percentage Composition

55. Determine the percentage composition of the following compounds:
    a. ammonium nitrate, NH_4NO_3, a common fertilizer
    b. tin(IV) oxide, SnO_2, an ingredient in fingernail polish

56. What percentage of ammonium carbonate, $(NH_4)_2CO_3$, an ingredient in smelling salts, is the ammonium ion, NH_4^+?

57. Some antacids use compounds of calcium, a mineral that is often lacking in the diet. What is the percentage composition of calcium carbonate, a common antacid ingredient?

MIXED REVIEW

58. Calculate the number of moles in each of the following samples:
    a. 8.2 g of sodium phosphate
    b. 6.66 g of calcium nitrate
    c. 8.22 g of sulfur dioxide

59. There are exactly 1000 mg in 1 g. A cup of hot chocolate has 35.0 mg of sodium ions, Na^+. One cup of milk has 290 mg of calcium ions, Ca^{2+}.

a. How many moles of sodium ions are in the cup of hot chocolate?

b. How many moles of calcium ions are in the milk?

60. Cyclopentane has the molecular formula C_5H_{10}. How many moles of hydrogen atoms are there in 4 moles of cyclopentane?

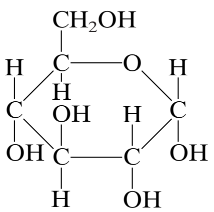
61. A 1.344 g sample of a compound contains 0.365 g Na, 0.221 g N, and 0.758 g O. What is its percentage composition? Calculate its empirical formula.

62. The naturally occurring silicon in sand has three isotopes; 92.23% is made up of atoms with a mass of 27.9769 amu, 4.67% is made up of atoms with a mass of 28.9765 amu, and 3.10% is made up of atoms with a mass of 29.9738 amu. Calculate the average atomic mass of silicon.

63. How many atoms of Fe are in the formula Fe_3C? How many moles of Fe are in one mole of Fe_3C?

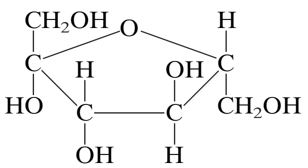
64. Shown below are the structures for two sugars, glucose and fructose.

a. What is the molar mass of glucose?



glucose

b. What is the molar mass of fructose?



fructose

65. Chlorine gas is a diatomic molecule, Cl_2. There are 6.00 mol of chlorine atoms in a sample of chlorine gas. How many moles of chlorine gas molecules is this?

66. Which yields a higher percentage of pure aluminum per gram, aluminum phosphate or aluminum chloride?

CRITICAL THINKING

67. Your calculation of the percentage composition of a compound gives 66.9% C and 29.6% H. Is the calculation correct? Explain.

68. Imagine you are a farmer, using NH_3 and NH_4NO_3 as sources of nitrogen. NH_3 costs \$0.50 per kg, and NH_4NO_3 costs \$0.25 per kg. Use percentage composition to decide which is the best buy for your money.

ALTERNATIVE ASSESSMENT

69. Research methods scientists initially used to find Avogadro's number. Then compare these methods with modern methods.

70. The most accurate method for determining the mass of an element involves a *mass spectrometer.* This instrument is also used to determine the isotopic composition of a natural element. Find out more about how a mass spectrometer works. Draw a model of how it works. Present the model to the class.

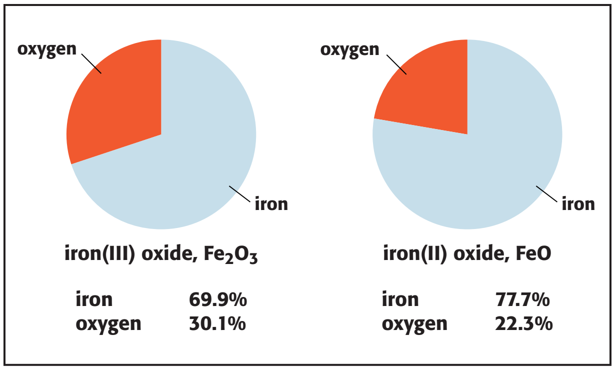
CONCEPT MAPPING

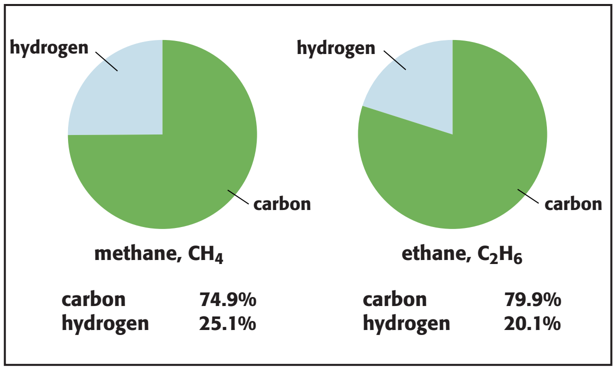
concept map

71. Use the following terms to create a concept map: *atoms, average atomic mass, molecules, mole, percentage composition,* and *molar masses*.

FOCUS ON GRAPHING

Study the graphs below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."





72. What do the slices of the pie represent?

73. What do the pie charts show about different compounds that are made up of the same elements?

74. Which has a higher percentage of oxygen, iron(II) oxide or iron(III) oxide?

75. Carlita has 30.0 g of oxygen and 70.0 g of iron. Can she make more FeO or Fe_2O_3 using only the reactants that she has?

76. a. Determine the percentage composition of propane, C_3H_8.

b. Make a pie chart for propane using a protractor to draw the correct sizes of the pie slices. (Hint: A circle has 360°. To draw the correct angle for each slice, multiply each percentage by 360°.)

c. Compare the charts for methane, ethane, and propane. How do the slices for carbon and hydrogen differ for each chart?

TECHNOLOGY AND LEARNING

77. Graphing Calculator

Calculating 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 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 MOLMASS and data sets and run the application as directed. If you are using another calculator, your teacher will provide you with the keystrokes and data sets to use. After you have graphed the data, answer the questions below.

a. What is the molar mass of $BaTiO_3$?

b. What is the molar mass of $PbCl_2$?

c. What is the molar mass of NH_4NO_3?

7

Standardized Test Prep

UNDERSTANDING CONCEPTS

***Directions (1–3):* For *each* question, write on a separate sheet of paper the letter of the correct answer.**

1. Element A has two isotopes. One has an atomic mass of 120 and constitutes 60%; the other has an atomic mass of 122 and constitutes 40%. Which range below includes the average atomic mass of Element A?
   A. less than 120
   B. between 120 and 121
   C. between 121 and 122
   D. greater than 122

2. Which of the following can be determined from the empirical formula of a compound alone?
   F. the true formula of the compound
   G. the molecular mass of the compound
   H. the percentage of composition of the compound
   I. the arrangement of atoms within a molecule of the compound

3. How many ions are in 0.5 moles of NaCl?
   A. 1.204×10^{23} **C.** 6.022×10^{23}
   B. 3.011×10^{23} **D.** 9.033×10^{23}

***Directions (4–6):* For *each* question, write a short response.**

4. How many moles of calcium (mass = 40.1) are in a serving of milk containing 290 mg of calcium?

5. How is Avogadro's number related to moles?

6. Antimony has two isotopes. One, amounting to 57.3% of the atoms, has a mass of 120.9. The other, 42.7% of the atoms, has a mass of 122.9. What is the average atomic mass of antimony?

READING SKILLS

***Directions (7–9):* Read the passage below. Then answer the questions.**

In 1800 two English chemists, Nicholson and Carlisle, discovered that when an electric current is passed through water, hydrogen and oxygen were produced in a 2:1 volume ratio and a 1:8 mass ratio. This evidence helped to support John Dalton's theory that matter consisted of atoms, demonstrating that water consists of the two elements in a constant proportion. If the same number of moles of each gas occupy the same volume, then each molecule of water must consist of twice as much hydrogen as oxygen, even though the mass of hydrogen is only one-eighth that of oxygen.

7. Based on this experiment, what is the empirical formula of water?
   F. HO
   G. H_2O
   H. H_2O_8
   I. HO_8

8. How would the experimental result have been different if hydrogen gas existed as individual atoms while oxygen formed molecules with two atoms bound by a covalent bond?
   A. The result would be the same.
   B. The ratio of hydrogen to oxygen would be 1:1.
   C. The ratio of hydrogen to oxygen would be 1:4.
   D. The ratio of hydrogen to oxygen would be 4:1.

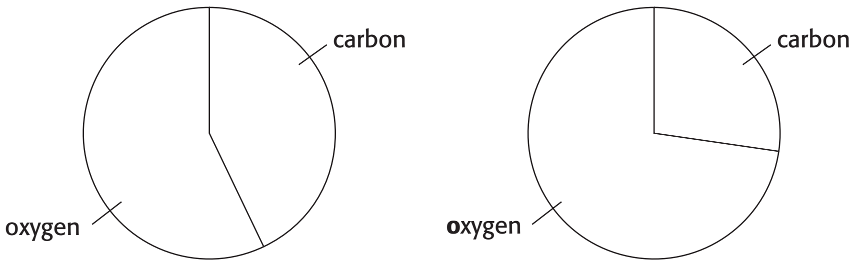
9. How does the empirical formula for water compare to its molecular formula?

INTERPRETING GRAPHICS

***Directions (10–13):* For *each* question below, record the correct answer on a separate sheet of paper.**

The charts below show the distribution of mass between carbon and oxygen in two compounds that are made of only those two elements. Use the diagram below to answer questions 10 through 13.

Carbon Monoxide and Carbon Dioxide



carbon monoxide, CO		**carbon dioxide, CO_2**	
carbon	42.88%	carbon	27.29%
oxygen	57.12%	oxygen	72.71%

10 How many moles of oxygen atoms are there in 100.0 moles of carbon dioxide?
F. 66.7
G. 72.7
H. 100.0
I. 200.0

11 Explain why the percentage of oxygen in carbon dioxide is **not** twice the percentage of oxygen in carbon monoxide, if there are twice as many oxygen atoms.

12 If you did not know the true formulas for carbon monoxide and carbon dioxide, what information would you need beyond what is provided in the illustration in order to calculate them?
A. the percentage compositions
B. the atomic masses of carbon and oxygen
C. the melting and boiling points of each compound
D. the number of atoms of each element in the compound

13 How many grams of carbon are contained in 200.0 grams of carbon dioxide?
F. 27.29
G. 42.88
H. 54.58
I. 85.76

Test *TIP*

When using a graph to answer a question, be sure to study the graph carefully before choosing a final answer. Some of the answer choices may be based on common misinterpretations of graphs.

CHAPTER 8

CHEMICAL EQUATIONS AND REACTIONS