Español

CHAPTER 11

The Mole

What You'll Learn

- You will use the mole and molar mass to make conversions among moles, mass, and number of representative particles.
- You will determine the percent composition of the components of compounds.
- You will calculate the empirical and molecular formulas for compounds and determine the formulas for hydrates.

Why It's Important

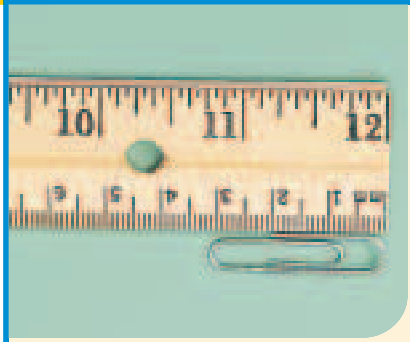
New materials, new products, new consumer goods of all kinds come on the market regularly. But before manufacturing begins on most new products, calculations involving the mole must be done.

Chemistry Online

Visit the Chemistry Web site at **chemistrymc.com** to find links to the mole.

Florists often sell flowers, such as roses, carnations, and tulips, by the dozen. A dozen is a counting unit for 12 items.

DISCOVERY LAB



Materials

centimeter ruler
paper clip

How much is a mole?

Counting large numbers of items is easier when you use counting units like the dozen. Chemists use a counting unit called the mole.

Procedure

1. Measure the length of a paper clip to the nearest 0.1 cm.
2. If a mole is 6.02×10^{23} items, how far will a mole of paper clips, placed end to end lengthwise, reach into space?

Analysis

How many light-years (ly) would the paper clips extend into space? (1 light-year $= 9.46 \times 10^{15}$ m). How does the distance you calculated compare with the following astronomical distances: nearest star (other than the sun) = 4.3 ly, center of our galaxy = 30 000 ly, nearest galaxy $= 2 \times 10^{6}$ ly?

Section 11.1 Measuring Matter

Objectives

- **Describe** how a mole is used in chemistry.
- **Relate** a mole to common counting units.
- **Convert** moles to number of representative particles and number of representative particles to moles.

Vocabulary

mole
Avogadro's number

If you were buying a bouquet of roses for a special occasion, you probably wouldn't ask for 12 or 24; you'd ask for one or two dozen. Similarly, you might buy a pair of gloves, a ream of paper for your printer, or a gross of pencils. Each of the units shown in **Figure 11-1**—a pair, a dozen, a gross, and a ream—represents a specific number of items. These units make counting objects easier. It's easier to buy and sell paper by the ream—500 sheets—than by the individual sheet.

Counting Particles

Each of the counting units shown in **Figure 11-1** is appropriate for certain kinds of objects depending primarily on their size and the use they serve. But regardless of the object—boots, eggs, pencils, paper—the number that the unit represents is always constant.

Figure 11-1

A pair is always two objects, a dozen is 12, a gross is 144, and a ream is 500. Can you think of any other counting units?

Chemists also need a convenient method for counting accurately the number of atoms, molecules, or formula units in a sample of a substance. As you know, atoms and molecules are extremely small. There are so many of them in even the smallest sample that it's impossible to actually count them. That's why chemists created their own counting unit called the mole. In the **DISCOVERY LAB**, you found that a mole of paper clips is an enormous number of items.

What is a mole? The **mole**, commonly abbreviated mol, is the SI base unit used to measure the amount of a substance. It is the number of representative particles, carbon atoms, in exactly 12 g of pure carbon-12. Through years of experimentation, it has been established that a mole of anything contains $6.022\ 136\ 7 \times 10^{23}$ representative particles. A representative particle is any kind of particle such as atoms, molecules, formula units, electrons, or ions. The number $6.022\ 136\ 7 \times 10^{23}$ is called **Avogadro's number** in honor of the Italian physicist and lawyer Amedeo Avogadro who, in 1811, determined the volume of one mole of a gas. In this book, Avogadro's number will be rounded to three significant figures—6.02×10^{23}.

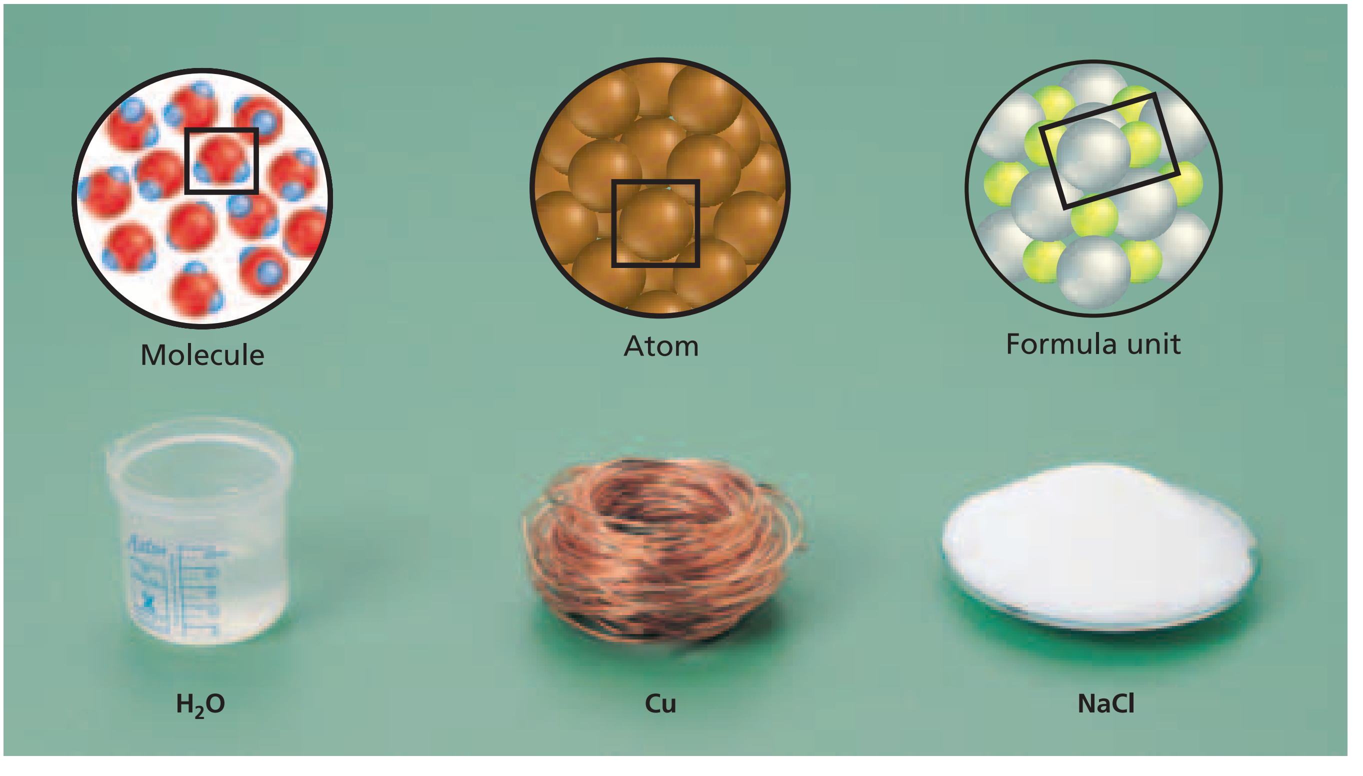
If you write out Avogadro's number, it looks like this.

$$602\ 000\ 000\ 000\ 000\ 000\ 000\ 000$$

Avogadro's number is an enormous number, as it must be in order to count extremely small particles. As you can imagine, Avogadro's number would not be convenient for measuring a quantity of marbles. Avogadro's number of marbles would cover the surface of Earth to a depth of more than six kilometers! But you can see in **Figure 11-2** that it is convenient to use the mole to measure substances. One-mole quantities of three substances are shown, each with a different representative particle. The representative particle in a mole of water is the water molecule, the representative particle in a mole of copper is the copper atom, and the representative particle in a mole of sodium chloride is the formula unit.

Figure 11-2

The amount of each substance shown is 6.02×10^{23} or one mole of representative particles. The representative particle for each substance is shown in a box. Refer to **Table C-1** in Appendix C for a key to atom color conventions.



Converting Moles to Particles and Particles to Moles

Suppose you buy three and a half dozen roses and want to know how many roses you have. Recall what you have learned about conversion factors. You can multiply the known quantity (3.5 dozen roses) by a conversion factor to express the quantity in the units you want (number of roses). You must set up your calculation as shown here so that all units cancel except those required for the answer.

$$\text{Conversion factor: } \frac{12 \text{ roses}}{1 \text{ dozen}}$$

$$3.5 \cancel{\text{dozen}} \times \frac{12 \text{ roses}}{1 \cancel{\text{dozen}}} = 42 \text{ roses}$$

Note that the units cancel and the answer tells you that 42 roses are in 3.5 dozen.

Now, suppose you want to determine how many particles of sucrose are in 3.50 moles of sucrose. You know that one mole contains 6.02×10^{23} representative particles. Therefore, you can write a conversion factor, Avogadro's number, that relates representative particles to moles of a substance.

$$\text{Conversion factor: } \frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mole}}$$

You can find the number of representative particles in a number of moles just as you found the number of roses in 3.5 dozen.

$$\text{number of moles} \times \frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mole}}$$
$$= \text{number of representative particles}$$

For sucrose, the representative particle is a molecule, so the number of molecules of sucrose is obtained by multiplying 3.50 moles of sucrose by the conversion factor, Avogadro's number.

$$3.50 \cancel{\text{mol sucrose}} \times \frac{6.02 \times 10^{23} \text{ molecules sucrose}}{1 \cancel{\text{mol sucrose}}}$$
$$= 2.11 \times 10^{24} \text{ molecules sucrose}$$

There are 2.11×10^{24} molecules of sucrose in 3.50 moles.

PRACTICE PROBLEMS

1. Determine the number of atoms in 2.50 mol Zn.

2. Given 3.25 mol $AgNO_3$, determine the number of formula units.

3. Calculate the number of molecules in 11.5 mol H_2O.

Now, suppose you want to find out how many moles are represented by a certain number of representative particles. You can use the inverse of Avogadro's number as a conversion factor.

$$\text{number of representative particles} \times \frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ representative particles}}$$
$$= \text{number of moles}$$

History CONNECTION

Lorenzo Romano Amedeo Carlo Avogadro, Conte di Quaregna e Ceretto was born in Turin, Italy in 1776 and was educated as a church lawyer. During the early 1800s, he studied mathematics and physics and was appointed to a professorship at the Royal College of Vercelli where he produced his hypothesis on gases. From 1820 until his death, Avogadro was professor of physics at the University of Turin where he conducted research on electricity and the physical properties of liquids.

Avogadro's hypothesis did not receive recognition for more than fifty years. Although Avogadro did nothing to measure the number of particles in equal volumes of gases, his hypothesis did lead to the eventual calculation of the number, 6.02×10^{23}.

National Mole Day is celebrated on October 23 (10/23) from 6:02 A.M. to 6:02 P.M. to commemorate Avogadro's contribution to modern chemistry.

For more practice converting from moles to representative particles, go to **Supplemental Practice Problems** in Appendix A.

The number of moles of substance is obtained by multiplying the number of particles by this factor, as you will see in Example Problem 11-1.

Ointments containing zinc oxide provide protection from sunburn and are used to treat some skin diseases.

EXAMPLE PROBLEM 11-1

Converting Number of Representative Particles to Moles

Zinc is used as a corrosion-resistant coating on iron and steel. It is also an essential trace element in your diet. Calculate the number of moles that contain 4.50×10^{24} atoms of zinc (Zn).

1. Analyze the Problem

You are given the number of atoms of zinc and must find the equivalent number of moles. If you compare 4.50×10^{24} atoms Zn with 6.02×10^{23}, the number of atoms in one mole, you can predict that the answer should be less than 10 moles.

Known	Unknown
number of atoms = 4.50×10^{24} atoms Zn	mol Zn = ? mol
1 mol Zn = 6.02×10^{23} atoms Zn	

2. Solve for the Unknown

Multiply the number of zinc atoms by the conversion factor that is the inverse of Avogadro's number.

$$\text{number of atoms} \times \frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ atoms}} = \text{number of moles}$$

$$4.50 \times 10^{24} \cancel{\text{atoms Zn}} \times \frac{1 \text{ mol Zn}}{6.02 \times 10^{23} \cancel{\text{atoms Zn}}} = 7.48 \text{ mol Zn}$$

3. Evaluate the Answer

The number of atoms of zinc and Avogadro's number have three significant figures. Therefore, the answer is expressed correctly with three digits. The answer is less than 10 moles, as predicted, and has the correct unit.

Practice!

For more practice converting from representative particles to moles, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

4. How many moles contain each of the following?

a. 5.75×10^{24} atoms Al

b. 3.75×10^{24} molecules CO_2

c. 3.58×10^{23} formula units $ZnCl_2$

d. 2.50×10^{20} atoms Fe

Section 11.1 Assessment

5. How is a mole similar to a dozen?

6. What is the relationship between Avogadro's number and one mole?

7. Explain how you can convert from the number of representative particles of a substance to moles of that substance.

8. Explain why chemists use the mole.

9. **Thinking Critically** Arrange the following from the smallest number of representative particles to the largest number of representative particles: 1.25×10^{25} atoms Zn; 3.56 mol Fe; 6.78×10^{22} molecules glucose ($C_6H_{12}O_6$).

10. **Using Numbers** Determine the number of representative particles in each of the following and identify the representative particle: 11.5 mol Ag; 18.0 mol H_2O; 0.150 mol NaCl.

Section 11.2 Mass and the Mole

You wouldn't expect a dozen limes to have the same mass as a dozen eggs. Eggs and limes differ in size and composition, so it's not surprising that they have different masses, as **Figure 11-3** shows. Moles of substances also have different masses for the same reason—the substances have different compositions. If you put a mole of carbon on a balance beside a mole of metallic copper, you would see a difference in mass just as you do for a dozen eggs and a dozen limes. Carbon atoms differ from copper atoms. Thus, the mass of 6.02×10^{23} atoms of carbon does not equal the mass of 6.02×10^{23} atoms of copper. How do you determine the mass of a mole?

Objectives

- **Relate** the mass of an atom to the mass of a mole of atoms.
- **Calculate** the number of moles in a given mass of an element and the mass of a given number of moles of an element.
- **Calculate** the number of moles of an element when given the number of atoms of the element.
- **Calculate** the number of atoms of an element when given the number of moles of the element.

Vocabulary

molar mass

Figure 11-3

A dozen limes has approximately twice the mass of a dozen eggs. The difference in mass is reasonable because limes are different from eggs in composition and size.

The Mass of a Mole

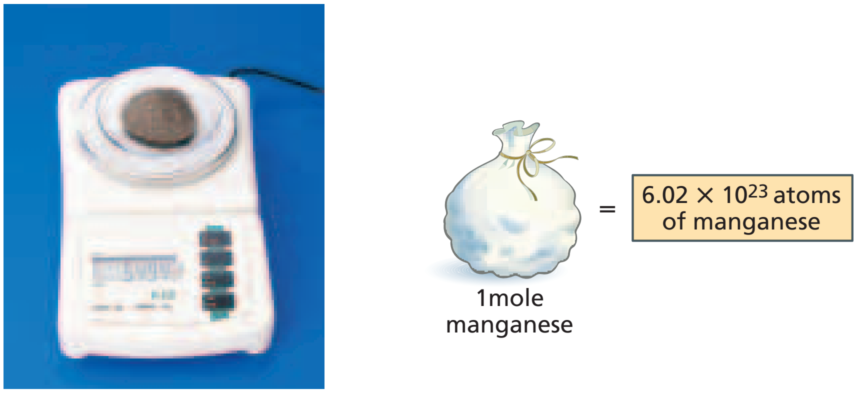
In Chapter 4, you learned that the relative scale of atomic masses uses the isotope carbon-12 as the standard. Each atom of carbon-12 has a mass of 12 atomic mass units (amu). The atomic masses of all other elements are established relative to carbon-12. For example, an atom of hydrogen-1 has a mass of 1 amu. The mass of an atom of helium-4 is 4 amu. Therefore, the mass of one atom of hydrogen-1 is one-twelfth the mass of one atom of carbon-12. The mass of one atom of helium-4 is one-third the mass of one atom of carbon-12.

You can find atomic masses on the periodic table, but notice that the values shown are not exact integers. For example, you'll find 12.011 amu for carbon, 1.008 amu for hydrogen, and 4.003 amu for helium. These differences occur because the recorded values are weighted averages of the masses of all the naturally occurring isotopes of each element.

How does the mass of one atom relate to the mass of a mole of that atom? You know that the mole is defined as the number of representative particles, or carbon-12 atoms, in exactly 12 g of pure carbon-12. Thus, the mass of one mole of carbon-12 atoms is 12 g. What about other elements? Whether you are considering a single atom or Avogadro's number of atoms (a mole), the masses of all atoms are established relative to the mass of carbon-12. The mass of a mole of hydrogen-1 is one-twelfth the mass of a mole of carbon-12 atoms, or 1.0 g. The mass of a mole of helium-4 atoms is one-third the mass of a mole of carbon-12 atoms, or 4.0 g. The mass in grams of one mole of any pure substance is called its **molar mass**. The molar mass of any element is numerically equal to its atomic mass and has the units g/mol. An atom of manganese has an atomic mass of 54.94 amu. Therefore, the molar mass of manganese is 54.94 g/mol. When you measure 54.94 g of manganese on a balance,

Figure 11-4

One mole of manganese, represented by a bag of particles, contains Avogadro's number of atoms and has a mass equal to its atomic mass in grams. The same is true for all the elements.



you indirectly count 6.02×10^{23} atoms of manganese. **Figure 11-4** shows the relationship between molar mass and one mole of an element. The **problem-solving LAB** will further clarify these relationships.

Using Molar Mass

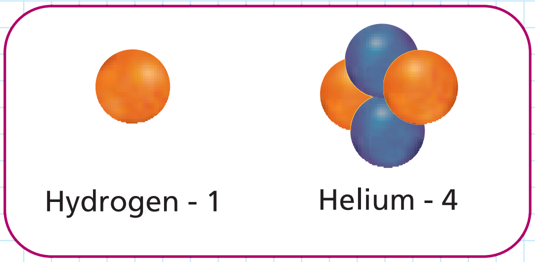
Imagine that your class bought jellybeans in bulk to sell by the dozen at a candy sale. You soon realize that it's too much work counting out each dozen, so instead you decide to measure the jellybeans by mass. You find that 1 dozen jellybeans has a mass of 35 g. What mass of jellybeans should you measure if a customer wants 5 dozen? The conversion factor that relates mass and dozens of jellybeans is

$$\frac{35 \text{ g jellybeans}}{1 \text{ dozen}}$$

problem-solving LAB

Molar Mass, Avogadro's Number and the Atomic Nucleus

Formulating models A nuclear model of mass can provide a simple picture of the connections between the mole, molar mass, and the number of representative particles in a mole.



Analysis

The diagram shows models of the nuclei of hydrogen-1 and helium-4. The hydrogen-1 nucleus contains one proton with a mass of 1.007 amu. The mass of the proton in grams has been found experimentally to be 1.672×10^{-24} g. Helium-4 contains two protons and two neutrons and has a mass of approximately 4 amu.

1. What is the mass in grams of one helium atom? (The mass of a neutron is approximately the same as the mass of a proton.)
2. Carbon-12 contains six protons and six neutrons. Draw a model of the nucleus of carbon-12 and calculate the mass of one atom in amu and in grams.

Thinking Critically

1. How many atoms of hydrogen-1 are in a 1.007-g sample? Recall that 1.007 amu is the mass of one atom of hydrogen-1. Round your answer to two significant digits.
2. If you had samples of helium and carbon that contained the same number of atoms as you calculated in question 1, what would be the mass in grams of each sample?
3. What can you conclude about the relationship between the number of atoms and the mass of each sample?

You would multiply the number of dozens to be sold by this conversion factor.

$$5\ \cancel{\text{dozen}} \times \frac{35\ \text{g jellybeans}}{1\ \cancel{\text{dozen}}} = 175\ \text{g jellybeans}$$

Note how the units cancel to give you the mass of 5 dozen jellybeans.

Now, suppose that while working in chemistry lab, you need 3.00 moles of manganese (Mn) for a chemical reaction. How can you measure that amount? Like the 5 dozen jellybeans, the number of moles of manganese can be converted to an equivalent mass and measured on a balance. To calculate mass from the number of moles, you need to multiply the number of moles of manganese required in the reaction (3.00 moles of Mn) by a conversion factor that relates mass and moles of manganese. That conversion factor is the molar mass of manganese (54.9 g/mol).

$$\text{number of moles} \times \frac{\text{number of grams}}{1\ \text{mole}} = \text{mass}$$

$$3.00\ \cancel{\text{mol Mn}} \times \frac{54.9\ \text{g Mn}}{1\ \cancel{\text{mol Mn}}} = 165\ \text{g Mn}$$

If you measure 165 g of manganese on a balance, you will have the 3.00 moles of manganese you need for the reaction. The reverse conversion—from mass to moles—also involves the molar mass as a conversion factor, but it is the inverse of the molar mass that is used. Can you explain why?

EXAMPLE PROBLEM 11-2

Mole to Mass Conversion

Chromium (Cr) is a transition element used as a coating on metals and in steel alloys to control corrosion. Calculate the mass in grams of 0.0450 moles of chromium.

1. Analyze the Problem

You are given the number of moles of chromium and must convert it to an equivalent mass using the molar mass of chromium from the periodic table. Because the sample is less than one-tenth mole, the answer should be less than one-tenth the molar mass.

Known	Unknown
number of moles = 0.0450 mol Cr	mass = ? g Cr
molar mass Cr = 52.00 g/mol Cr	

2. Solve for the Unknown

Multiply the known number of moles of chromium by the conversion factor that relates grams of chromium to moles of chromium, the molar mass.

$$\text{moles Cr} \times \frac{\text{grams Cr}}{1\ \text{mol Cr}} = \text{grams Cr}$$

$$0.0450\ \cancel{\text{mol Cr}} \times \frac{52.00\ \text{g Cr}}{1\ \cancel{\text{mol Cr}}} = 2.34\ \text{g Cr}$$

3. Evaluate the Answer

The known number of moles of chromium has the smallest number of significant figures (3), so the answer is correctly stated with three digits. The answer is less than one-tenth the mass of one mole as predicted and has the correct unit.

Chromium resists corrosion, which means it doesn't react readily with oxygen in the air. It was used in this 1948 Cadillac to protect the steel and add glitter.

PRACTICE PROBLEMS

11. Determine the mass in grams of each of the following.

a. 3.57 mol Al
b. 42.6 mol Si
c. 3.45 mol Co
d. 2.45 mol Zn

EXAMPLE PROBLEM 11-3

Mass to Mole Conversion

Calcium, the fifth most abundant element on Earth, is always found combined with other elements because of its high reactivity. How many moles of calcium are in 525 g calcium (Ca)?

1. Analyze the Problem

You are given the mass of calcium and must convert the mass to moles of calcium. The mass of calcium is more than ten times larger than the molar mass. Therefore, the answer should be greater than ten moles.

Known	Unknown
mass = 525 g Ca	number of moles = ? mol Ca
molar mass Ca = 40.08 g/mol Ca	

2. Solve for the Unknown

Multiply the known amount of calcium by the conversion factor that relates moles of calcium to grams of calcium, the inverse of molar mass.

$$\text{mass} \times \frac{1\ \text{mole}}{\text{number of grams}} = \text{number of moles}$$

$$525\ \cancel{\text{g Ca}} \times \frac{1\ \text{mol Ca}}{40.08\ \cancel{\text{g Ca}}} = 13.1\ \text{mol Ca}$$

3. Evaluate the Answer

The mass of calcium has the smaller number of significant figures (3), so the answer is expressed correctly with three digits. As predicted, the answer is greater than 10 moles and has the expected unit.

Math Handbook

Review the meaning of inverse in the **Math Handbook** on page 905 of this text.

Practice!

For more practice with mass and mole conversions, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

12. Determine the number of moles in each of the following.

a. 25.5 g Ag
b. 300.0 g S
c. 125 g Zn
d. 1.00 kg Fe

Conversions from mass to atoms and atoms to mass So far, you have learned how to convert mass to the number of moles and the number of moles to mass. You can go one step further and convert mass to the number of atoms. Recall the jellybeans you were selling at the candy sale. At the end of the day, you find that 550 g of jellybeans are left unsold. Without counting, can you determine how many jellybeans this is? You know that one dozen jellybeans has a mass of 35 g and that 1 dozen contains 12 jellybeans. Thus, you can first convert the 550 g to dozens of jellybeans by using the conversion factor that relates dozens and mass.

$$550\ \cancel{\text{g jellybeans}} \times \frac{1\ \text{dozen jellybeans}}{35\ \cancel{\text{g jellybeans}}} = 16\ \text{dozen jellybeans}$$

Next, you can determine how many jellybeans are in 16 dozen by multiplying by the conversion factor that relates number of particles (jellybeans) and dozens.

$$16 \cancel{\text{dozen}} \times \frac{12 \text{ jellybeans}}{1 \cancel{\text{dozen}}} = 192 \text{ jellybeans}$$

The 550 g of leftover jellybeans is equal to 192 jellybeans.

Just as you cannot make a direct conversion from the mass of jellybeans to the number of jellybeans, you cannot make a direct conversion from the mass of a substance to the number of representative particles in that substance. You must first convert the mass to moles by multiplying by a conversion factor that relates moles and mass. Can you identify the conversion factor? The number of moles must then be multiplied by a conversion factor that relates the number of representative particles to moles. That conversion factor is Avogadro's number.

EXAMPLE PROBLEM 11-4

Mass to Atoms Conversion

Gold is one of a group of metals called the coinage metals (copper, silver, and gold). How many atoms of gold (Au) are in a pure gold nugget having a mass of 25.0 g.

1. Analyze the Problem

You are given a mass of gold and must determine how many atoms it contains. Because you cannot go directly from mass to the number of atoms, you must first convert mass to moles using molar mass. Then, you can convert moles to the number of atoms using Avogadro's number. The given mass of the gold nugget is about one-eighth the molar mass of gold (196.97 g/mol), so the number of gold atoms should be approximately one-eighth Avogadro's number.

Known	Unknown
mass = 25.0 g Au	number of atoms = ? atoms Au
molar mass Au = 196.97 g/mol Au	

2. Solve for the Unknown

Multiply the known amount of gold by the inverse of the molar mass as the conversion factor.

$$\text{mass Au} \times \frac{1 \text{ mole Au}}{\text{number of grams Au}} = \text{moles Au}$$

$$25.0 \cancel{\text{g Au}} \times \frac{1 \text{ mol Au}}{196.97 \cancel{\text{g Au}}} = 0.127 \text{ mol Au}$$

Multiply the calculated number of moles of gold by Avogadro's number as a conversion factor.

$$\text{moles Au} \times \frac{6.02 \times 10^{23} \text{ atoms Au}}{1 \text{ mole Au}} = \text{atoms Au}$$

$$0.127 \cancel{\text{mol Au}} \times \frac{6.02 \times 10^{23} \text{ atoms Au}}{1 \cancel{\text{mol Au}}} = 7.65 \times 10^{22} \text{ atoms Au}$$

3. Evaluate the Answer

The mass of gold has the smallest number of significant figures (3), so the answer is expressed correctly with three digits. The answer is approximately one-eighth Avogadro's number as predicted, and the unit is correct.

Gold is called a noble metal because it doesn't react readily with other elements. Early civilizations used nearly pure gold for coins and ornaments such as this gold mask from Quimbaya, Columbia, A.D. 1000-1500.

PRACTICE PROBLEMS

13. How many atoms are in each of the following samples?

a. 55.2 g Li
b. 0.230 g Pb
c. 11.5 g Hg
d. 45.6 g Si
e. 0.120 kg Ti

EXAMPLE PROBLEM 11-5

Atoms to Mass Conversion

Helium is an unreactive noble gas often found in underground deposits mixed with methane. The mixture is separated by cooling the gaseous mixture until all but the helium has liquified.

A party balloon contains 5.50×10^{22} atoms of helium (He) gas. What is the mass in grams of the helium?

1. Analyze the Problem

You are given the number of atoms of helium and must find the mass of the gas.

Known	Unknown
number of atoms = 5.50×10^{22} atoms He	mass = ? g He
molar mass He = 4.00 g/mol He	

2. Solve for the Unknown

Multiply the number of atoms of helium by the inverse of Avogadro's number as a conversion factor.

$$\cancel{\text{atoms He}} \times \frac{1 \text{ mol He}}{6.02 \times 10^{23} \cancel{\text{ atoms He}}} = \text{moles He}$$

$$5.50 \times 10^{22} \cancel{\text{ atoms He}} \times \frac{1 \text{ mol He}}{6.02 \times 10^{23} \cancel{\text{ atoms He}}} = 0.0914 \text{ mol He}$$

Multiply the calculated number of moles of helium by the conversion factor that relates mass of helium to moles of helium, molar mass.

$$\text{moles He} \times \frac{\text{number of grams He}}{1 \text{ mole He}} = \text{mass He}$$

$$0.0914 \cancel{\text{ mol He}} \times \frac{4.00 \text{ g He}}{1 \cancel{\text{ mol He}}} = 0.366 \text{ g He}$$

3. Evaluate the Answer

The answer is expressed correctly with three significant figures and has the expected unit.

Helium gas, used in party balloons, is heavier than hydrogen gas but safer because it is unreactive and will not burn as hydrogen does.

Practice!

For more practice with mass and number of atoms conversions, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

14. What is the mass in grams of each of the following?

a. 6.02×10^{24} atoms Bi
b. 1.00×10^{24} atoms Mn
c. 3.40×10^{22} atoms He
d. 1.50×10^{15} atoms N
e. 1.50×10^{15} atoms U

Now that you have learned about and practiced conversions between mass, moles, and representative particles, you can see that the mole is at the center of these calculations. Mass must always be converted to moles before being converted to atoms, and atoms must similarly be converted to moles before calculating their mass. **Figure 11-5** shows the steps to follow as you work with these conversions.

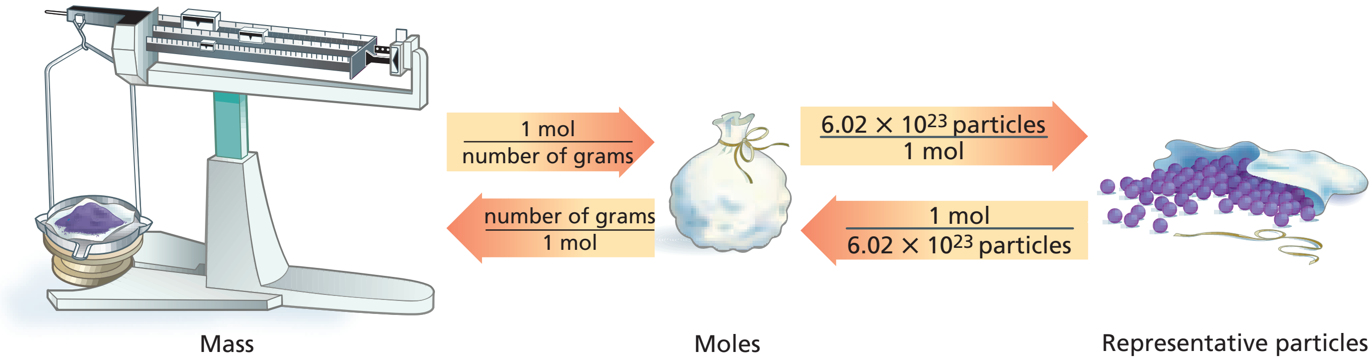


Figure 11-5

The mole is at the center of conversions between mass and particles. Two steps are needed to go from mass to representative particles or the reverse.

In **Figure 11-5**, mass is represented by a laboratory balance, moles are represented by a bag or bundle of particles, and representative particles are represented by the contents that are spilling out of the bag. You can see that two steps are needed to convert from mass on the left to representative particles on the right or to convert from representative particles on the right to mass on the left. The conversion factors for these conversions are given on the arrows pointing left and right. In the Example Problems, you have been making each of these conversions in separate steps, but you could make the same conversions in one calculation. For example, suppose you want to find out how many molecules of water are in 1.00 g of water. This calculation involves the conversion factors on the arrows pointing to the right. You could set up your calculation like this.

$$1.00\ \cancel{g\ H_2O} \times \frac{1\ \cancel{mol\ H_2O}}{18.02\ \cancel{g\ H_2O}} \times \frac{6.02 \times 10^{23}\ \text{molecules}\ H_2O}{1\ \cancel{mol\ H_2O}}$$

$$= 3.34 \times 10^{22}\ \text{molecules}\ H_2O$$

Note that the units cancel to give the answer in molecules of water. Do the reverse calculation yourself using the conversion factors on the arrows pointing from right to left. What is the mass of 3.34×10^{22} molecules of water? What answer should you expect? What unit?

Section 11.2 Assessment

15. Explain what is meant by molar mass.

16. What conversion factor should be used to convert from mass to moles? Moles to mass?

17. Explain the steps needed to convert the mass of an element to the number of atoms of the element.

18. Thinking Critically The mass of a single atom is usually given in the unit amu. Would it be possible to express the mass of a single atom in grams? Explain.

19. Sequencing Arrange the following in order of mass from the smallest mass to the largest: 1.0 mol Ar, 3.0×10^{24} atoms Ne, 20 g Kr.

Section 11.3 Moles of Compounds

Español

Objectives

- **Recognize** the mole relationships shown by a chemical formula.
- **Calculate** the molar mass of a compound.
- **Calculate** the number of moles of a compound from a given mass of the compound, and the mass of a compound from a given number of moles of the compound.
- **Determine** the number of atoms or ions in a mass of a compound.

You have learned that different kinds of representative particles are counted using the mole, but so far you have applied this counting unit only to atoms of elements. Can you make similar conversions for compounds and ions? If so, you will need to know the molar mass of the compounds and ions.

Chemical Formulas and the Mole

Recall that the chemical formula for a compound indicates the types of atoms and the number of each contained in one unit of the compound. For example, freon has the formula CCl_2F_2. The subscripts in the formula tell you that one molecule of CCl_2F_2 consists of one atom of carbon, two atoms of chlorine, and two atoms of fluorine that have chemically combined. The ratio of carbon to chlorine to fluorine is 1 : 2 : 2.

But suppose you had a mole of freon. The representative particles would be molecules of freon. A mole of freon would contain Avogadro's number of freon molecules, which means that instead of one carbon atom, you would have a mole of carbon atoms. And instead of two chlorine atoms and two fluorine atoms, you would have two moles of chlorine atoms and two moles of fluorine atoms. The ratio of carbon to chlorine to fluorine in one mole of freon would still be 1 : 2 : 2, as it is in one molecule of freon.

Figure 11-6 illustrates this principle for a dozen freon molecules. Check for yourself that a dozen freon molecules contains one dozen carbon atoms, two dozen chlorine atoms, and two dozen fluorine atoms. The chemical formula CCl_2F_2 not only represents an individual molecule of freon, it also represents a mole of the compound.

In some chemical calculations, you may need to convert from moles of a compound to moles of individual atoms in the compound or from moles of individual atoms in a compound to moles of the compound. The following conversion factors can be written for use in these calculations for the molecule freon.

$$\frac{1 \text{ mol C atoms}}{1 \text{ mol } CCl_2F_2} \qquad \frac{2 \text{ mol Cl atoms}}{1 \text{ mol } CCl_2F_2} \qquad \frac{2 \text{ mol F atoms}}{1 \text{ mol } CCl_2F_2}$$

To find out how many moles of fluorine atoms are in 5.50 moles of freon, you would multiply the moles of freon by the conversion factor that relates moles of fluorine atoms to moles of CCl_2F_2.

$$\text{moles } CCl_2F_2 \times \frac{\text{moles F atoms}}{1 \text{ mole } CCl_2F_2} = \text{moles F atoms}$$

$$5.50 \cancel{\text{mol } CCl_2F_2} \times \frac{2 \text{ mol F atoms}}{1 \cancel{\text{mol } CCl_2F_2}} = 11.0 \text{ mol F atoms}$$

Therefore, 11.0 mol F atoms are in 5.50 mol CCl_2F_2.

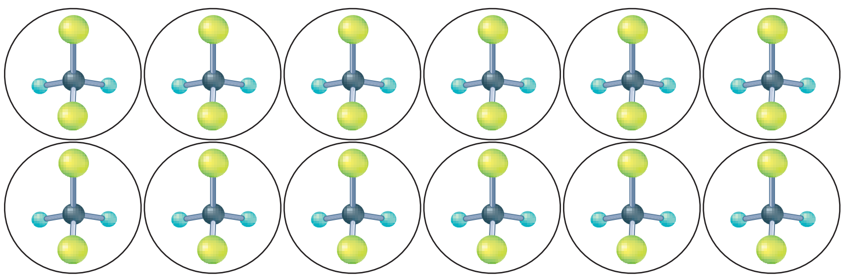


Figure 11-6

A dozen freon molecules contains one dozen carbon atoms, two dozen chlorine atoms, and two dozen fluorine atoms. How many of each kind of atom are contained in one mole of freon?

Conversion factors such as the one just used for fluorine can be written for any element in a compound. The number of moles of the element that goes in the numerator of the conversion factor is the subscript for that element in the chemical formula.

EXAMPLE PROBLEM 11-6

Mole Relationships from a Chemical Formula

Aluminum oxide (Al_2O_3), often called alumina, is the principal raw material for the production of aluminum. Alumina occurs in the minerals corundum and bauxite. Determine the moles of aluminum ions (Al^{3+}) in 1.25 moles of aluminum oxide.

1. Analyze the Problem

You are given the number of moles of Al_2O_3 and must determine the number of moles of Al^{3+} ions. Use a conversion factor based on the chemical formula that relates moles of Al^{3+} ions to moles of Al_2O_3. Every mole of Al_2O_3 contains two moles of Al^{3+} ions. Thus, the answer should be two times the number of moles of Al_2O_3.

Known

number of moles = 1.25 mol Al_2O_3

Unknown

number of moles = ? mol Al^{3+} ions

2. Solve for the Unknown

1 mol Al_2O_3 contains 2 mol Al^{3+} ions. Determine the conversion factor relating moles of Al^{3+} ions to moles of Al_2O_3.

$$\frac{2 \text{ mol } Al^{3+} \text{ ions}}{1 \text{ mol } Al_2O_3}$$

Multiply the known number of moles of Al_2O_3 by the conversion factor.

$$\text{moles } Al_2O_3 \times \frac{\text{moles } Al^{3+} \text{ ions}}{\text{mole } Al_2O_3} = \text{moles } Al^{3+} \text{ ions}$$

$$1.25 \cancel{\text{ mol } Al_2O_3} \times \frac{2 \text{ mol } Al^{3+} \text{ ions}}{1 \cancel{\text{ mol } Al_2O_3}} = 2.50 \text{ mol } Al^{3+} \text{ ions}$$

3. Evaluate the Answer

Because the conversion factor is a ratio of whole numbers, the number of significant digits is based on the moles of Al_2O_3. Therefore, the answer is expressed correctly with three significant figures. As predicted, the answer is twice the number of moles of Al_2O_3.

The capstone placed at the top of the Washington Monument in 1884 is a 22.86-cm pyramid of pure aluminum. Until an inexpensive purification process was developed, aluminum was considered a rare and precious metal.

PRACTICE PROBLEMS

20. Determine the number of moles of chloride ions in 2.50 mol $ZnCl_2$.

21. Calculate the number of moles of each element in 1.25 mol glucose ($C_6H_{12}O_6$).

22. Determine the number of moles of sulfate ions present in 3.00 mol iron(III) sulfate ($Fe_2(SO_4)_3$).

23. How many moles of oxygen atoms are present in 5.00 mol diphosphorus pentoxide (P_2O_5)?

24. Calculate the number of moles of hydrogen atoms in 11.5 mol water.

For more practice calculating the number of moles of atoms or ions in a given number of moles of a compound, go to **Supplemental Practice Problems** in Appendix A.

The Molar Mass of Compounds

The mass of your backpack is the sum of the mass of the pack plus the masses of the books, notebooks, pencils, lunch, and miscellaneous items you put into it. You could find its mass by determining the mass of each item separately and adding them together. Similarly, the mass of a mole of a compound equals the sum of the masses of every particle that makes up the compound. You know how to use the molar mass of an element as a conversion factor in calculations. You also know that a chemical formula indicates the number of moles of each element in a compound. With this information, you can now determine the molar mass of a compound.

Suppose you want to determine the molar mass of potassium chromate (K_2CrO_4). Using the periodic table, the mass of one mole of each element present in potassium chromate can be determined. That mass is then multiplied by the number of moles of that element in the chemical formula. Adding the masses of all elements present will yield the molar mass of K_2CrO_4.

$$\text{number of moles} \times \text{molar mass} = \text{number of grams}$$

$$2.000\ \cancel{\text{mol K}} \times \frac{39.10\text{ g K}}{1\ \cancel{\text{mol K}}} = 78.20\text{ g}$$

$$1.000\ \cancel{\text{mol Cr}} \times \frac{52.00\text{ g Cr}}{1\ \cancel{\text{mol Cr}}} = 52.00\text{ g}$$

$$4.000\ \cancel{\text{mol O}} \times \frac{16.00\text{ g O}}{1\ \cancel{\text{mol O}}} = \underline{64.00\text{ g}}$$

$$\text{molar mass } K_2CrO_4 = 194.20\text{ g}$$

PRACTICE PROBLEMS

25. Determine the molar mass of each of the following ionic compounds: $NaOH$, $CaCl_2$, $KC_2H_3O_2$, $Sr(NO_3)_2$, and $(NH_4)_3PO_4$.

26. Calculate the molar mass of each of the following molecular compounds: C_2H_5OH, $C_{12}H_{22}O_{11}$, HCN, CCl_4, and H_2O.

Practice!

For more practice calculating the molar mass of a compound, go to **Supplemental Practice Problems** in Appendix A.

The molar mass of a compound demonstrates the law of conservation of mass. The sum of the masses of the elements that reacted to form the compound equals the mass of the compound. **Figure 11-7** shows 194 g, or one mole, of K_2CrO_4 and masses equal to one mole of two other substances.

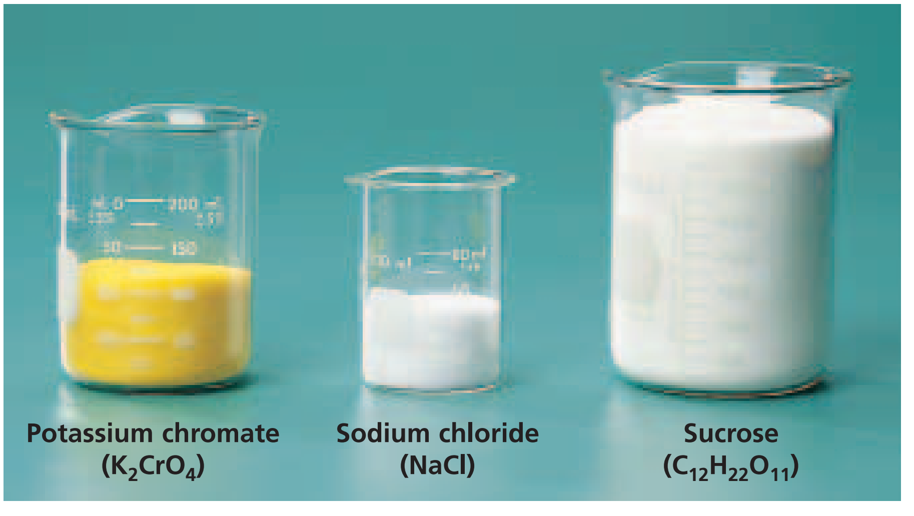


Figure 11-7

Each substance contains different numbers and kinds of atoms so their molar masses are different. The molar mass of each compound is the sum of the masses of all the elements contained in the compound.

Converting Moles of a Compound to Mass

Suppose you need to measure a certain number of moles of a compound for an experiment. First, you must calculate the mass in grams that corresponds to the necessary number of moles. Then, that mass can be measured on a balance. In Example Problem 11-2, you learned how to convert the number of moles of elements to mass using molar mass as the conversion factor. The procedure is the same for compounds except that you must first calculate the molar mass of the compound.

EXAMPLE PROBLEM 11-7

Mole-to-Mass Conversion for Compounds

The characteristic odor of garlic is due to the compound allyl sulfide ($(C_3H_5)_2S$). What is the mass of 2.50 moles of allyl sulfide?

The pungent odor of garlic is characteristic of sulfides. Sulfides, including hydrogen sulfide, are noted for their strong, often unpleasant odors. The sulfur atom in allyl sulfide forms a chemical bond to each of the two C_3H_5 groups in the molecule.

1. Analyze the Problem

You are given 2.50 mol $(C_3H_5)_2S$ and must convert the moles to mass using the molar mass as a conversion factor. The molar mass is the sum of the molar masses of all the elements in $(C_3H_5)_2S$.

Known

number of moles = 2.50 mol $(C_3H_5)_2S$

Unknown

molar mass $(C_3H_5)_2S$ = ? g/mol $(C_3H_5)_2S$

mass = ? g $(C_3H_5)_2S$

2. Solve for the Unknown

Calculate the molar mass of $(C_3H_5)_2S$.

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

$$6\ \cancel{\text{mol C}} \times \frac{12.01\text{ g C}}{1\ \cancel{\text{mol C}}} = 72.06\text{ g C}$$

$$10\ \cancel{\text{mol H}} \times \frac{1.008\text{ g H}}{1\ \cancel{\text{mol H}}} = \underline{10.08\text{ g H}}$$

$$\text{molar mass } (C_3H_5)_2S = 114.21\text{ g/mol } (C_3H_5)_2S$$

Convert mol $(C_3H_5)_2S$ to g $(C_3H_5)_2S$ by using the molar mass as a conversion factor.

$$\text{moles } (C_3H_5)_2S \times \frac{\text{number of grams } (C_3H_5)_2S}{1\text{ mole } (C_3H_5)_2S} = \text{mass } (C_3H_5)_2S$$

$$2.50\ \cancel{\text{mol } (C_3H_5)_2S} \times \frac{114.21\text{ g } (C_3H_5)_2S}{1\ \cancel{\text{mol } (C_3H_5)_2S}} = 286\text{ g } (C_3H_5)_2S$$

3. Evaluate the Answer

Mol $(C_3H_5)_2S$ has the smaller number of significant figures (3), so the answer is expressed correctly with three digits. The unit, g, is correct.

PRACTICE PROBLEMS

27. What is the mass of 3.25 moles of sulfuric acid (H_2SO_4)?

28. What is the mass of 4.35×10^{-2} moles of zinc chloride ($ZnCl_2$)?

29. How many grams of potassium permanganate are in 2.55 moles?

For more practice converting moles of a compound to mass, go to **Supplemental Practice Problems** in Appendix A.

Converting the Mass of a Compound to Moles

Imagine that the experiment you are doing in the laboratory produces 5.55 g of a compound. How many moles is this? To find out, you calculate the molar mass of the compound and determine it to be 185.0 g/mol. The molar mass relates grams and moles, but this time you need the inverse of the molar mass as the conversion factor.

$$5.50\ \cancel{\text{g compound}} \times \frac{1 \text{ mol compound}}{185.0\ \cancel{\text{g compound}}} = 0.0297 \text{ mol compound}$$

EXAMPLE PROBLEM 11-8

This compound, commonly called lime, is calcium oxide (CaO). Calcium oxide reacts with water to produce calcium hydroxide. Lime is a component of cement and is used to counteract excess acidity in soil.

Mass-to-Mole Conversion for Compounds

Calcium hydroxide ($Ca(OH)_2$) is used to remove sulfur dioxide from the exhaust gases emitted by power plants and for softening water by the elimination of Ca^{2+} and Mg^{2+} ions. Calculate the number of moles of calcium hydroxide in 325 g.

1. Analyze the Problem

You are given 325 g $Ca(OH)_2$ and are solving for the number of moles of $Ca(OH)_2$. You must first calculate the molar mass of $Ca(OH)_2$.

Known

mass = 325 g $Ca(OH)_2$

Unknown

molar mass = ? g/mol $Ca(OH)_2$

number of moles = ? mol $Ca(OH)_2$

2. Solve for the Unknown

Determine the molar mass of $Ca(OH)_2$.

$$1 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} = 40.08 \text{ g}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{mol H}} = \underline{2.016 \text{ g}}$$

$$\text{molar mass of } Ca(OH)_2 = 74.096 \text{ g/mol} = 74.10 \text{ g/mol}$$

Use the inverse of molar mass as the conversion factor to calculate moles.

$$325\ \cancel{\text{g } Ca(OH)_2} \times \frac{1 \text{ mol } Ca(OH)_2}{74.10\ \cancel{\text{g } Ca(OH)_2}} = 4.39 \text{ mol } Ca(OH)_2$$

3. Evaluate the Answer

The given mass of $Ca(OH)_2$ has fewer digits than any other value in the calculations so it determines the number of significant figures in the answer (3). To check the reasonableness of the answer, round off the molar mass of $Ca(OH)_2$ to 75 g/mol and the given mass of $Ca(OH)_2$ to 300 g. Seventy- five is contained in 300 four times. Thus, the answer is reasonable.

PRACTICE PROBLEMS

30. Determine the number of moles present in each of the following.

a. 22.6 g $AgNO_3$

b. 6.50 g $ZnSO_4$

c. 35.0 g HCl

d. 25.0 g Fe_2O_3

e. 254 g $PbCl_4$

Practice!

For more practice converting the mass of a compound to moles, go to **Supplemental Practice Problems** in Appendix A.

Converting the Mass of a Compound to Number of Particles

Example Problem 11-8 illustrated how to find the number of moles of a compound contained in a given mass. Now, you will learn how to calculate the number of representative particles—molecules or formula units—contained in a given mass and, in addition, the number of atoms or ions. Recall that no direct conversion is possible between mass and number of particles. You must first convert the given mass to moles by multiplying by the inverse of the molar mass. Then, you can convert moles to the number of representative particles by multiplying by Avogadro's number. To determine numbers of atoms or ions in a compound, you will need conversion factors that are ratios of the number of atoms or ions in the compound to one mole of compound. These are based on the chemical formula. Example Problem 11-9 provides practice in solving this type of problem.

EXAMPLE PROBLEM 11-9

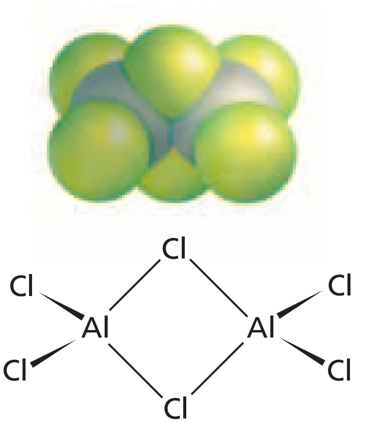
Conversion from Mass to Moles to Particles

Aluminum chloride is used in refining petroleum and manufacturing rubber and lubricants. A sample of aluminum chloride ($AlCl_3$) has a mass of 35.6 g.

a. How many aluminum ions are present?

b. How many chloride ions are present?

c. What is the mass in grams of one formula unit of aluminum chloride?



At ordinary temperatures, aluminum chloride is a solid with the formula $AlCl_3$. In the vapor phase, however, aluminum chloride exists as a doubled molecule, or dimer, with the formula Al_2Cl_6.

1. Analyze the Problem

You are given 35.6 g $AlCl_3$ and must calculate the number of Al^{3+} ions, the number of Cl^- ions, and the mass in grams of one formula unit of $AlCl_3$. Molar mass, Avogadro's number, and ratios from the chemical formula are the necessary conversion factors. The ratio of Al^{3+} ions to Cl^- ions in the chemical formula is 1:3. Therefore, the calculated numbers of ions should be in that ratio. The mass of one formula unit in grams should be an extremely small number.

Known	Unknown
mass = 35.6 g $AlCl_3$	number of ions = ? Al^{3+} ions
	number of ions = ? Cl^- ions
	mass = ? g/formula unit $AlCl_3$

2. Solve for the Unknown

Determine the molar mass of $AlCl_3$.

$$1\ \cancel{\text{mol Al}} \times \frac{26.98\text{ g Al}}{1\ \cancel{\text{mol Al}}} = 26.98\text{ g Al}$$

$$3\ \cancel{\text{mol Cl}} \times \frac{35.45\text{ g Cl}}{1\ \cancel{\text{mol Cl}}} = \underline{106.35\text{ g Cl}}$$

$$\text{Molar mass of AlCl}_3 = 133.33\text{ g/mol AlCl}_3$$

Multiply by the inverse of the molar mass as a conversion factor to convert the mass of $AlCl_3$ to moles.

$$\cancel{\text{grams AlCl}_3} \times \frac{1\text{ mol AlCl}_3}{\cancel{\text{grams AlCl}_3}} = \text{moles AlCl}_3$$

$$35.6\ \cancel{\text{g AlCl}_3} \times \frac{1\text{ mol AlCl}_3}{133.33\ \cancel{\text{g AlCl}_3}} = 0.267\text{ mol AlCl}_3$$

Continued on next page

Multiply by Avogadro's number to calculate the number of formula units of $AlCl_3$.

$$0.267\ \cancel{mol\ AlCl_3} \times \frac{6.02 \times 10^{23}\ \text{formula units}}{1\ \cancel{mol\ AlCl_3}} = 1.61 \times 10^{23}\ \text{formula units}\ AlCl_3$$

To calculate the number of Al^{3+} and Cl^- ions, use the ratios from the chemical formula as conversion factors.

$$1.61 \times 10^{23}\ \cancel{AlCl_3\ formula\ unit} \times \frac{1\ Al^{3+}\ \text{ion}}{1\ \cancel{AlCl_3\ formula\ unit}} = 1.61 \times 10^{23}\ Al^{3+}\ \text{ions}$$

$$1.61 \times 10^{23}\ \cancel{AlCl_3\ formula\ unit} \times \frac{3\ Cl^-\ \text{ions}}{1\ \cancel{AlCl_3\ formula\ unit}} = 4.83 \times 10^{23}\ Cl^-\ \text{ions}$$

Calculate the mass in grams of one formula unit of $AlCl_3$. Start with molar mass and use the inverse of Avogadro's number as a conversion factor.

$$\frac{133.33\ g\ AlCl_3}{1\ \cancel{mol}} \times \frac{1\ \cancel{mol}}{6.02 \times 10^{23}\ \text{formula unit}} = 2.21 \times 10^{-22}\ g\ AlCl_3/\text{formula unit}$$

3. Evaluate the Answer

A minimum of three significant figures is used in each value in the calculations. Therefore, the answers have the correct number of digits. The number of Cl^- ions is three times the number of Al^{3+} ions, as predicted. The mass of a formula unit of $AlCl_3$ can be checked by calculating it in a different way: Divide the mass of $AlCl_3$ (35.6 g) by the number of formula units contained in the mass (1.61×10^{23} formula units) to obtain the mass of one formula unit. The two answers are the same.

PRACTICE PROBLEMS

31. A sample of silver chromate (Ag_2CrO_4) has a mass of 25.8 g.

a. How many Ag^+ ions are present?

b. How many CrO_4^{2-} ions are present?

c. What is the mass in grams of one formula unit of silver chromate?

32. What mass of sodium chloride contains 4.59×10^{24} formula units?

33. A sample of ethanol (C_2H_5OH) has a mass of 45.6 g.

a. How many carbon atoms does the sample contain?

b. How many hydrogen atoms are present?

c. How many oxygen atoms are present?

34. A sample of sodium sulfite (Na_2SO_3) has a mass of 2.25 g.

a. How many Na^+ ions are present?

b. How many SO_3^{2-} ions are present?

c. What is the mass in grams of one formula unit of Na_2SO_3?

35. A sample of carbon dioxide has a mass of 52.0 g.

a. How many carbon atoms are present?

b. How many oxygen atoms are present?

c. What is the mass in grams of one molecule of CO_2?

For more practice calculating the number of ions or atoms in a mass of a compound and the mass in grams of one formula unit, go to **Supplemental Practice Problems** in Appendix A.

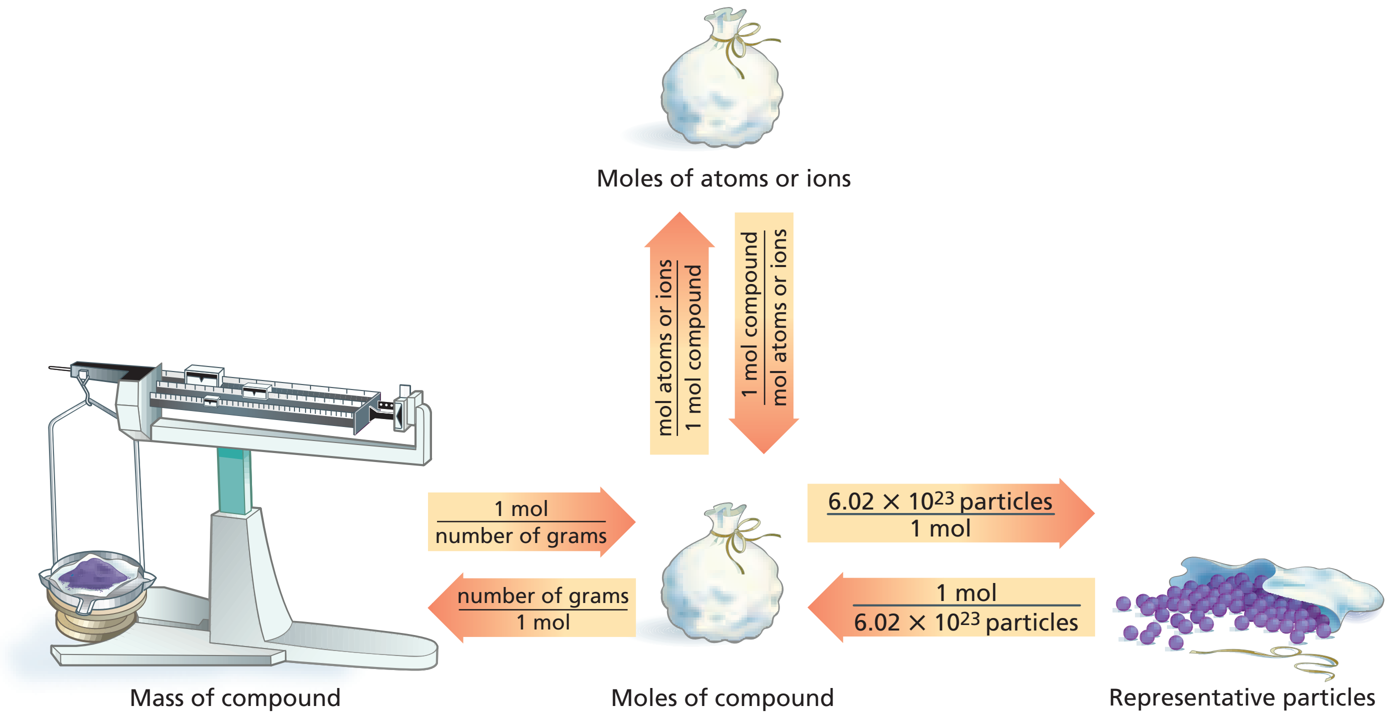


Figure 11-8

Note the central position of the mole. To go from the left, right, or top of the diagram to any other place, you must go through the mole. The conversion factors on the arrows provide the means for making the conversions.

Conversions among mass, moles, and the number of particles are summarized in **Figure 11-8**. Refer to this diagram often until you become familiar with the calculations. Note that the molar mass (number of grams/1 mol) and the inverse of molar mass (1 mol/number of grams) are the conversion factors between the mass of a substance and the number of moles of the substance. Avogadro's number and its inverse are the conversion factors between the moles of a substance and the number of representative particles. To convert between the number of moles of a compound and the number of moles of atoms or ions contained in the compound, you need the ratio of moles of atoms or ions to 1 mole of compound or its inverse, which are shown on the upward and downward arrows in **Figure 11-8**. These ratios are derived from the subscripts in the chemical formula. What ratio would you use to find the moles of hydrogen atoms in four moles of water?

Section 11.3 Assessment

36. Describe how you can determine the molar mass of a compound.

37. What three conversion factors are often used in mole conversions?

38. Explain how you can determine the number of atoms or ions in a given mass of a compound.

39. If you know the mass in grams of a molecule of a substance, could you obtain the mass of a mole of that substance? Explain.

40. **Thinking Critically** Design a bar graph that will show the number of moles of each element present in 500 g dioxin ($C_{12}H_4Cl_4O_2$), a powerful poison.

41. **Applying Concepts** The recommended daily allowance of calcium is 1000 mg of Ca^{2+} ions. Calcium carbonate is used to supply the calcium in vitamin tablets. How many moles of calcium ions does 1000 mg represent? How many moles of calcium carbonate are needed to supply the required amount of calcium ions? What mass of calcium carbonate must each tablet contain?

Section 11.4

Empirical and Molecular Formulas

Español

Objectives

- **Explain** what is meant by the percent composition of a compound.
- **Determine** the empirical and molecular formulas for a compound from mass percent and actual mass data.

Vocabulary

percent composition
empirical formula
molecular formula

Chemists, such as those shown in **Figure 11-9**, are often involved in developing new compounds for industrial, pharmaceutical, and home uses. After a synthetic chemist (one who makes new compounds) has produced a new compound, an analytical chemist analyzes the compound to provide experimental proof of its composition and its chemical formula. You can learn more about the work of chemists by reading **Chemistry and Technology** at the end of this chapter.

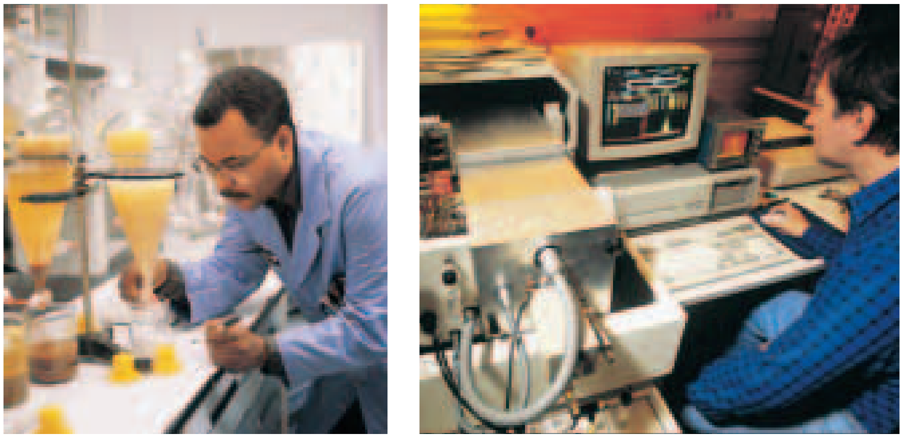


Figure 11-9

New compounds are first made on a small scale by a synthetic chemist like the one on the left. Then, an analytical chemist, like the one on the right, analyzes the compound to verify its structure and percent composition.

Percent Composition

It's the analytical chemist's job to identify the elements a compound contains and determine their percent by mass. Gravimetric and volumetric analyses are experimental procedures based on the measurement of mass for solids and liquids, respectively. For example, a 100-g sample of a new compound contains 55 g of element X and 45 g of element Y. The percent by mass of any element in a compound can be found by dividing the mass of the element by the mass of the compound and multiplying by 100.

$$\frac{\text{mass of element}}{\text{mass of compound}} \times 100 = \text{percent by mass}$$

$$\frac{55 \text{ g element X}}{100 \text{ g compound}} \times 100 = 55\% \text{ element X}$$

$$\frac{45 \text{ g element Y}}{100 \text{ g compound}} \times 100 = 45\% \text{ element Y}$$

Because percent means parts per 100, the percents by mass of all the elements of a compound must always add up to 100. The percent composition of the compound is 55% X and 45% Y. The percent by mass of each element in a compound is called the **percent composition** of a compound.

Percent composition from the chemical formula If you already know the chemical formula for a compound such as water (H_2O), can you calculate its percent composition? The answer is yes. You can use the chemical formula to calculate the molar mass of water (18.02 g/mol) and assume you have an 18.02-g sample. Because the percent composition of a compound is always the same, no matter the size of the sample, you can assume that the sample size is one mole. To find the mass of each element in a mole of water, multiply the molar mass of the element by its subscript in the chemical formula. Because one mole of water contains two moles of hydrogen atoms, the mass of hydrogen in a mole of water is (2 mol)(1.01 g/mol) = 2.02 g. To find the percent by mass of hydrogen in water, divide the mass of hydrogen by the molar mass of water (18.02 g/mol) and multiply by 100.

$$\frac{2.02 \text{ g H}}{18.02 \text{ g H}_2\text{O}} \times 100 = 11.2\% \text{ H}$$

One mole of water contains one mole of oxygen. Thus, the mass of oxygen in one mole of water is 16.00 g. The percent by mass of oxygen is

$$\frac{16.00 \text{ g O}}{18.02 \text{ g H}_2\text{O}} \times 100 = 88.80\% \text{ O}$$

The percent composition of water is 11.2% hydrogen and 88.80% oxygen.

The general equation for calculating the percent by mass of any element in a compound is

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

The **miniLAB** provides an opportunity to practice calculating percents.

Careers Using Chemistry

Analytical Chemist

Would you like to work in a laboratory, solving problems and making precise measurements with state-of-the-art equipment? If so, consider a career as an analytical chemist.

Analytical chemists identify and measure the elements and compounds found in substances. They might determine the composition of the raw materials used in manufacturing or help physicians diagnose diseases. They might identify air or water pollutants or determine which nutrients are in certain foods. Analytical chemists work in fields as varied as archeology, crime, and space science.

miniLAB

Percent Composition and Gum

Interpreting Data Water soluble sweeteners and flavorings are added to chewing gum. Are these chemicals added as an outside coating or are they mixed throughout the gum?

Materials balance, weighing paper, 250-mL beakers (2), pieces of chewing gum (2), stirring rod, paper towels, window screen (10 cm × 10 cm), scissors, clock or timer

Procedure

CAUTION: *Do not taste or eat any items used in the lab.*

1. Unwrap two pieces of chewing gum. Measure the mass of each separately on a piece of weighing paper. Label the weighing papers with the masses to avoid mixing up your data. Record the masses.
2. Add 150 mL of cold tap water to a 250-mL beaker. Place one piece of chewing gum in the water and stir for two minutes.
3. Remove the gum from the water and pat dry using paper towels. Measure and record the mass of the dried gum.
4. Use scissors to cut the second piece of gum into small pieces, each about the width of a pea. Repeat step 2 using fresh water. Use the stirring rod to keep the pieces of gum from clumping together.
5. Use the window screen to strain the water from the gum. Pat the gum dry using paper towels. Measure and record the mass of the dried gum.
6. Discard the gum in a waste container.

Analysis

1. For the uncut piece of gum, calculate the mass of sweeteners and flavorings that dissolved in the water. The mass of sweeteners and flavorings is the difference between the original mass of the gum and the mass of the dried gum.
2. For the gum that was in small pieces, calculate the mass of dissolved sweeteners and flavorings.
3. For both pieces of gum, calculate the percent of the original mass that was soluble sweeteners and flavorings. For help, refer to *Percents* in the **Math Handbook** on page 909 of this text.
4. What can you infer from the two percentages? Is the gum sugar-coated or are the sweeteners and flavorings mixed throughout?

EXAMPLE PROBLEM 11-10

Calculating Percent Composition

Sodium hydrogen carbonate, also called baking soda, is an active ingredient in some antacids used for the relief of indigestion. Determine the percent composition of sodium hydrogen carbonate ($NaHCO_3$).

1. Analyze the Problem

You are given only the chemical formula. Assume you have one mole of $NaHCO_3$. Calculate the molar mass and the mass of each element in one mole to determine the percent by mass of each element in the compound. The sum of all percents should be 100%.

Known	Unknown
formula = $NaHCO_3$	percent Na = ? % Na
	percent H = ? % H
	percent C = ? % C
	percent O = ? % O

2. Solve for the Unknown

Determine the mass of each element present and the molar mass of $NaHCO_3$.

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

$$1\ \cancel{\text{mol H}} \times \frac{1.008\text{ g H}}{1\ \cancel{\text{mol H}}} = 1.008\text{ g H}$$

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

$$3\ \cancel{\text{mol O}} \times \frac{16.00\text{ g O}}{1\ \cancel{\text{mol O}}} = \underline{48.00\text{ g O}}$$

$$\text{molar mass } NaHCO_3 = 84.008\text{ g/mol } NaHCO_3 = 84.01\text{g/mol } NaHCO_3$$

Determine the percent by mass of each element by dividing the mass of the element by the molar mass of the compound and multiplying by 100.

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

$$\text{percent Na} = \frac{22.99\text{ g Na}}{84.01\text{ g } NaHCO_3} \times 100 = 27.37\%\text{ Na}$$

$$\text{percent H} = \frac{1.008\text{ g H}}{84.01\text{ g } NaHCO_3} \times 100 = 1.200\%\text{ H}$$

$$\text{percent C} = \frac{12.01\text{ g C}}{84.01\text{ g } NaHCO_3} \times 100 = 14.30\%\text{ C}$$

$$\text{percent O} = \frac{48.00\text{ g O}}{84.01\text{ g } NaHCO_3} \times 100 = 57.14\%\text{ O}$$

The percent composition of $NaHCO_3$ is 27.37% Na, 1.200% H, 14.30% C, and 57.14% O.

3. Evaluate the Answer

All masses and molar masses contain four significant figures. Therefore, the percents are correctly stated to four signigicant figures. The sum of the mass percents is 100.00% as required.

Try at Home LAB

See page 957 in Appendix E for **Calculating Carbon Percentages**

Antacids often contain carbonates, for example, sodium hydrogen carbonate, calcium carbonate, and magnesium carbonate. A carbonate-containing antacid neutralizes excess stomach acid by reacting with acid to produce carbon dioxide.

PRACTICE PROBLEMS

42. Determine the percent by mass of each element in calcium chloride.

43. Calculate the percent composition of sodium sulfate.

44. Which has the larger percent by mass of sulfur, H_2SO_3 or $H_2S_2O_8$?

45. What is the percent composition of phosphoric acid (H_3PO_4)?

Practice!

For more practice calculating percent composition, go to **Supplemental Practice Problems** in Appendix A.

Empirical Formula

Suppose that the identities of the elements in a sample of a new compound have been determined and the compound's percent composition is known. These data can be used to find the formula for the compound. First, you must determine the smallest whole number ratio of the moles of the elements in the compound. This ratio provides the subscripts in the empirical formula. The **empirical formula** for a compound is the formula with the smallest whole-number mole ratio of the elements. The empirical formula may or may not be the same as the actual molecular formula. If the two formulas are different, the molecular formula will always be a simple multiple of the empirical formula. The empirical formula for hydrogen peroxide is HO; the molecular formula is H_2O_2. In both formulas, the ratio of oxygen to hydrogen is 1:1.

The data used to determine the chemical formula for a compound may be in the form of percent composition or it may be the actual masses of the elements in a given mass of the compound. If percent composition is given, you can assume that the total mass of the compound is 100.00 g and that the percent by mass of each element is equal to the mass of that element in grams. For example, the percent composition of an oxide of sulfur is 40.05% S and 59.95% O. Thus, as you can see in **Figure 11-10**, 100.00 g of the oxide contains 40.05 g S and 59.95 g O. The mass of each element can be converted to a number of moles by multiplying by the inverse of the molar mass. Recall that the number of moles of S and O are calculated in this way.

$$40.05\ \cancel{\text{g S}} \times \frac{1\ \text{mol S}}{32.07\ \cancel{\text{g S}}} = 1.249\ \text{mol S}$$

$$59.95\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 3.747\ \text{mol O}$$

The mole ratio of S atoms to O atoms in the oxide is 1.249 : 3.747. As you can see, these values are not whole numbers and cannot be used as subscripts in a chemical formula.

How, then, can the mole ratio be converted to whole numbers? As a starting point, recognize that the element with the smaller number of moles, in this case sulfur, might have the smallest subscript possible, 1. You can make the mole value of sulfur equal to 1 if you divide both mole values by the value of sulfur (1.249). In doing so, you do not change the ratio between the two elements because both are divided by the same number.

$$\frac{1.249\ \text{mol S}}{1.25} = 1\ \text{mol S}$$

$$\frac{3.747\ \text{mol O}}{1.25} = 3\ \text{mol O}$$

The simplest whole number mole ratio of S atoms to O atoms is 1:3. Thus, the empirical formula for the oxide of sulfur is SO_3.

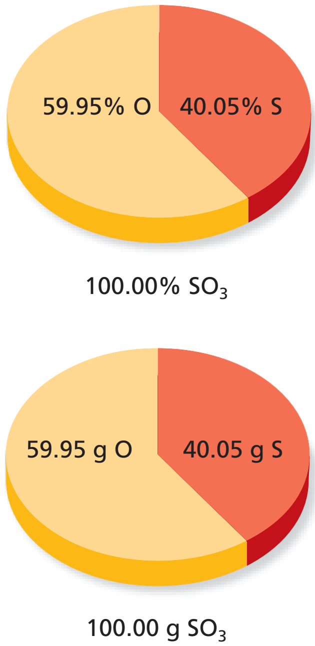


Figure 11-10

Keep this figure in mind when doing problems using percent composition. You can always assume that you have a 100-g sample of the compound and use the percents of the elements as masses of the elements.

Often in determining empirical formulas, the calculated mole values are still not whole numbers, as they are in the preceding example. In such cases, all the mole values must be multiplied by the smallest factor that will make them whole numbers. This is shown in Example Problem 11-11.

Math Handbook

Review ratios in the **Math Handbook** on page 908 of this text.

EXAMPLE PROBLEM 11-11

Calculating an Empirical Formula from Percent Composition

Methyl acetate is a solvent commonly used in some paints, inks, and adhesives. Determine the empirical formula for methyl acetate, which has the following chemical analysis: 48.64% carbon, 8.16% hydrogen, and 43.20% oxygen.

1. Analyze the Problem

You are given the percent composition of methyl acetate and must find the empirical formula. Because you can assume that each percent by mass represents the mass of the element in a 100.00-g sample, the percent sign can be replaced with the unit grams. Then, you can convert from grams to moles using the molar mass and find the smallest whole-number ratio of moles of the elements.

Known	Unknown
percent by mass = 48.64% C	empirical formula = ?
percent by mass = 8.16% H	
percent by mass = 43.20% O	

2. Solve for the Unknown

The mass of C is 48.64 g, the mass of H is 8.16 g, and the mass of O is 43.20 g. Multiply the mass of each element by the conversion factor that relates moles to grams based on molar mass.

$$48.64\,\cancel{\text{g C}} \times \frac{1\text{ mol C}}{12.01\,\cancel{\text{g C}}} = 4.050\text{ mol C}$$

$$8.16\,\cancel{\text{g H}} \times \frac{1\text{ mol H}}{1.008\,\cancel{\text{g H}}} = 8.10\text{ mol H}$$

$$43.20\,\cancel{\text{g O}} \times \frac{1\text{ mol O}}{16.00\,\cancel{\text{g O}}} = 2.700\text{ mol O}$$

Methyl acetate has a mole ratio of 4.050 mol C : 8.10 mol H : 2.700 mol O.

Calculate the simplest ratio of moles of the elements by dividing each number of moles by the smallest value in the mole ratio.

$$\frac{4.050\text{ mol C}}{2.700} = 1.500\text{ mol C} = 1.5\text{ mol C}$$

$$\frac{8.10\text{ mol H}}{2.700} = 3.00\text{ mol H} = 3\text{ mol H}$$

$$\frac{2.700\text{ mol O}}{2.700} = 1.000\text{ mol O} = 1\text{ mol O}$$

The simplest ratio is 1.5 mol C : 3 mol H : 1 mol O.

Multiply the numbers of moles in the ratio by the smallest number that will produce a ratio of whole numbers.

$2 \times 1.5\text{ mol C} = 3\text{ mol C}$

$2 \times 3\text{ mol H} = 6\text{ mol H}$

$2 \times 1\text{ mol O} = 2\text{ mol O}$

The simplest whole-number ratio of C atoms to H atoms to O atoms is 3 : 6 : 2. The empirical formula is $C_3H_6O_2$.

CONTENTS

3. Evaluate the Answer

The calculations are correct and significant figures have been observed. To check that the formula is correct, the percent composition represented by the formula can be calculated. The percent composition checks exactly with the data given in the problem.

PRACTICE PROBLEMS

46. A blue solid is found to contain 36.84% nitrogen and 63.16% oxygen. What is the empirical formula for this solid?

47. Determine the empirical formula for a compound that contains 35.98% aluminum and 64.02% sulfur.

48. Propane is a hydrocarbon, a compound composed only of carbon and hydrogen. It is 81.82% carbon and 18.18% hydrogen. What is the empirical formula?

49. The chemical analysis of aspirin indicates that the molecule is 60.00% carbon, 4.44% hydrogen, and 35.56% oxygen. Determine the empirical formula for aspirin.

50. What is the empirical formula for a compound that contains 10.89% magnesium, 31.77% chlorine, and 57.34% oxygen?

Practice!

For more practice calculating an empirical formula from percent composition, go to **Supplemental Practice Problems** in Appendix A.

Molecular Formula

Would it surprise you to learn that two or more substances with distinctly different properties can have the same percent composition and the same empirical formula? How is this possible? Remember that the subscripts in an empirical formula indicate the simplest whole-number ratio of moles of the elements in the compound. But the simplest ratio does not always indicate the actual number of moles in the compound. To identify a new compound, a chemist must go one step further and determine the **molecular formula**, which specifies the actual number of atoms of each element in one molecule or formula unit of the substance. **Figure 11-11** shows an important use of the gas, acetylene. It has the same percent composition and empirical formula, CH, as benzene which is a liquid. Yet chemically and structurally acetylene and benzene are very different.

To determine the molecular formula for a compound, the molar mass of the compound must be determined through experimentation and compared with the mass represented by the empirical formula. For example, the molar mass of acetylene is 26.04 g/mol and the mass of the empirical formula, CH, is13.02 g/mol. Dividing the actual molar mass by the mass of the empirical formula indicates that the molar mass of acetylene is two times the mass of the empirical formula.

$$\frac{\text{experimentally determined molar mass of acetylene}}{\text{mass of empirical formula CH}} =$$

$$\frac{26.04\ \cancel{\text{g/mol}}}{13.02\ \cancel{\text{g/mol}}} = 2.000$$

The result shows that the molar mass of acetylene is two times the mass represented by the empirical formula. Thus, the molecular formula of acetylene must contain twice the number of carbon atoms and twice the number of hydrogen atoms represented by the empirical formula.

Figure 11-11

Acetylene is a gas used for welding because of the high-temperature flame produced when it is burned with oxygen.

Chemistry Online

Topic: Empirical Formulas
To learn more about empirical formulas, visit the Chemistry Web site at **chemistrymc.com**

Activity: Research how empirical formulas are used in mineral identification. Prepare a chart comparing the formulas of the minerals with their properties.

Similarly, when the experimentally determined molar mass of benzene, 78.12 g/mol, is compared with the mass of the empirical formula, the molar mass of benzene is found to be six times the mass of the empirical formula.

$$\frac{\text{experimentally determined molar mass of benzene}}{\text{mass of empirical formula CH}} =$$

$$\frac{78.12\ \cancel{\text{g/mol}}}{13.02\ \cancel{\text{g/mol}}} = 6.000$$

The molar mass of benzene is six times the mass represented by the empirical formula, so the molecular formula for benzene must represent six times the number of carbon and hydrogen atoms shown in the empirical formula. You can conclude that the molecular formula for acetylene is (CH)2 or C_2H_2 and the molecular formula for benzene is (CH)6 or C_6H_6.

A molecular formula can be represented as the empirical formula multiplied by an integer *n*.

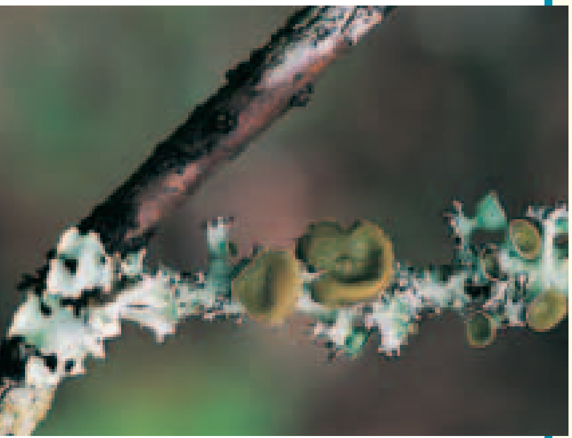
$$\text{molecular formula} = (\text{empirical formula})n$$

The integer is the factor (6 in the example above) by which the subscripts in the empirical formula must be multiplied to obtain the molecular formula.

EXAMPLE PROBLEM 11-12

Determining a Molecular Formula

Succinic acid is a substance produced by lichens. Chemical analysis indicates it is composed of 40.68% carbon, 5.08% hydrogen, and 54.24% oxygen and has a molar mass of 118.1 g/mol. Determine the empirical and molecular formulas for succinic acid.



Succinic acid occurs naturally in fossils and fungi, and in lichens such as those shown. Succinic acid produced commercially is used to make compounds used in perfumes (esters) and in lacquers and dyes.

1. Analyze the Problem

You are given the percent composition that allows you to calculate the empirical formula. Assume that each percent by mass represents the mass of the element in a 100.00-g sample.You can compare the given molar mass with the mass represented by the empirical formula to find *n*.

Known	Unknown
percent by mass = 40.68% C	empirical formula = ?
percent by mass = 5.08% H	molecular formula = ?
percent by mass = 54.24 % O	
molar mass = 118.1 g/mol succinic acid	

2. Solve for the Unknown

Use the percents by mass as grams of elements and convert to the number of moles by multiplying by the conversion factor that relates moles to mass based on molar mass.

$$40.68\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 3.387\ \text{mol C}$$

$$5.08\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}} = 5.04\ \text{mol H}$$

$$54.24\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 3.390\ \text{mol O}$$

Succinic acid has a mole ratio of C : H : O of 3.387 : 5.04 : 3.390.

Calculate the simplest ratio among the elements by dividing each mole value by the smallest value in the mole ratio.

$$\frac{3.387 \text{ mol C}}{3.387} = 1 \text{ mol C}$$

$$\frac{5.040 \text{ mol H}}{3.387} = 1.49 \text{ mol H} = 1.5 \text{ mol H}$$

$$\frac{3.390 \text{ mol O}}{3.387} = 1.001 \text{ mol O} = 1 \text{ mol O}$$

The simplest ratio is 1 mol C : 1.5 mol H : 1 mol O.

The simplest mol ratio includes a fractional value that cannot be used as a subscript in a formula. Multiply all mole values by 2.

$2 \times 1 \text{ mol C} = 2 \text{ mol C}$

$2 \times 1.5 \text{ mol H} = 3 \text{ mol H}$

$2 \times 1 \text{ mol O} = 2 \text{ mol O}$

The simplest whole-number ratio of C atoms to H atoms to O atoms is 2 : 3 : 2. The empirical formula is $C_2H_3O_2$

Calculate the empirical formula mass using the molar mass of each element.

$$2 \cancel{\text{mol C}} \times \frac{12.01 \text{ g C}}{1 \cancel{\text{mol C}}} = 24.02 \text{ g C}$$

$$3 \cancel{\text{mol H}} \times \frac{1.008 \text{ g H}}{1 \cancel{\text{mol H}}} = 3.024 \text{ g H}$$

$$2 \cancel{\text{mol O}} \times \frac{16.00 \text{ g O}}{1 \cancel{\text{mol O}}} = \underline{32.00 \text{ g O}}$$

$$\text{molar mass } C_2H_3O_2 = 59.04 \text{ g/mol } C_2H_3O_2$$

Divide the experimentally determined molar mass of succinic acid by the mass of the empirical formula to determine *n*.

$$n = \frac{\text{molar mass of succinic acid}}{\text{molar mass of } C_2H_3O_2}$$

$$n = \frac{118.1 \cancel{\text{g/mol}}}{59.04 \cancel{\text{g/mol}}} = 2.000$$

Multiply the subscripts in the empirical formula by 2 to determine the actual subscripts in the molecular formula.

$(C_2H_3O_2)(2) = C_4H_6O_4$

The molecular formula for succinic acid is $C_4H_6O_4$.

3. Evaluate the Answer

Calculation of the molar mass from the molecular formula gives the same result as the experimental molar mass.

PRACTICE PROBLEMS

51. Analysis of a chemical used in photographic developing fluid indicates a chemical composition of 65.45% C, 5.45% H, and 29.09% O. The molar mass is found to be 110.0 g/mol. Determine the molecular formula.

52. A compound was found to contain 49.98 g carbon and 10.47 g hydrogen. The molar mass of the compound is 58.12 g/mol. Determine the molecular formula.

53. A colorless liquid composed of 46.68% nitrogen and 53.32% oxygen has a molar mass of 60.01 g/mol. What is the molecular formula?

Practice!

For more practice calculating a molecular formula from percent composition, go to **Supplemental Practice Problems** in Appendix A.

EXAMPLE PROBLEM 11-13

Because titanium is stable under severe conditions of heat and cold, it is used in aircraft engines, missiles, and space vehicles.

Calculating an Empirical Formula from Mass Data

Although the mineral ilmenite contains more iron than titanium, the ore is usually mined and processed for titanium, a strong, light, and flexible metal. A sample of ilmenite is found to contain 5.41 g iron, 4.64 g titanium, and 4.65 g oxygen. Determine the empirical formula for ilmenite.

1. Analyze the Problem

You are given the masses of the elements found in a known mass of ilmenite and must determine the empirical formula of the mineral. Convert the known masses of each element to moles using the conversion factor that relates moles to grams, the inverse of molar mass. Then, find the smallest whole-number ratio of the moles of the elements.

Known

mass of iron = 5.41 g Fe
mass of titanium = 4.64 g Ti
mass of oxygen = 4.65 g O

Unknown

empirical formula = ?

2. Solve for the Unknown

Multiply the known mass of each element by the conversion factor that relates moles to grams based on the molar mass.

$$5.41\ \cancel{\text{g Fe}} \times \frac{1\text{ mol Fe}}{55.85\ \cancel{\text{g Fe}}} = 0.0969\text{ mol Fe}$$

$$4.64\ \cancel{\text{g Ti}} \times \frac{1\text{ mol Ti}}{47.88\ \cancel{\text{g Ti}}} = 0.0969\text{ mol Ti}$$

$$4.65\ \cancel{\text{g O}} \times \frac{1\text{ mol O}}{16.00\ \cancel{\text{g O}}} = 0.291\text{ mol O}$$

Ilmenite has a mole ratio of Fe : Ti : O of 0.0969 : 0.0969 : 0.291. Calculate the simplest ratio by dividing each mol value by the smallest value in the ratio. Fe and Ti have the same lower value (0.0969).

$$\frac{0.0969\text{ mol Fe}}{0.0969} = 1\text{ mol Fe}$$

$$\frac{0.0969\text{ mol Ti}}{0.0969} = 1\text{ mol Ti}$$

$$\frac{0.291\text{ mol O}}{0.0969} = 3\text{ mol O}$$

All the mol values are whole numbers. Thus, the simplest whole-number ratio of Fe : Ti : O is 1 : 1 : 3. The empirical formula for ilmenite is $FeTiO_3$.

3. Evaluate the Answer

The mass of iron is slightly greater than the mass of titanium, but the molar mass of iron is also slightly greater than that of titanium. Thus, it is reasonable that the number of moles of iron and titanium are equal. The mass of titanium is approximately the same as the mass of oxygen, but the molar mass of oxygen is about 1/3 that of titanium. Thus, a 3:1 ratio of oxygen to titanium is reasonable.

PRACTICE PROBLEMS

54. When an oxide of potassium is decomposed, 19.55 g K and 4.00 g O are obtained. What is the empirical formula for the compound?

55. Analysis of a compound composed of iron and oxygen yields 174.86 g Fe and 75.14 g O. What is the empirical formula for this compound?

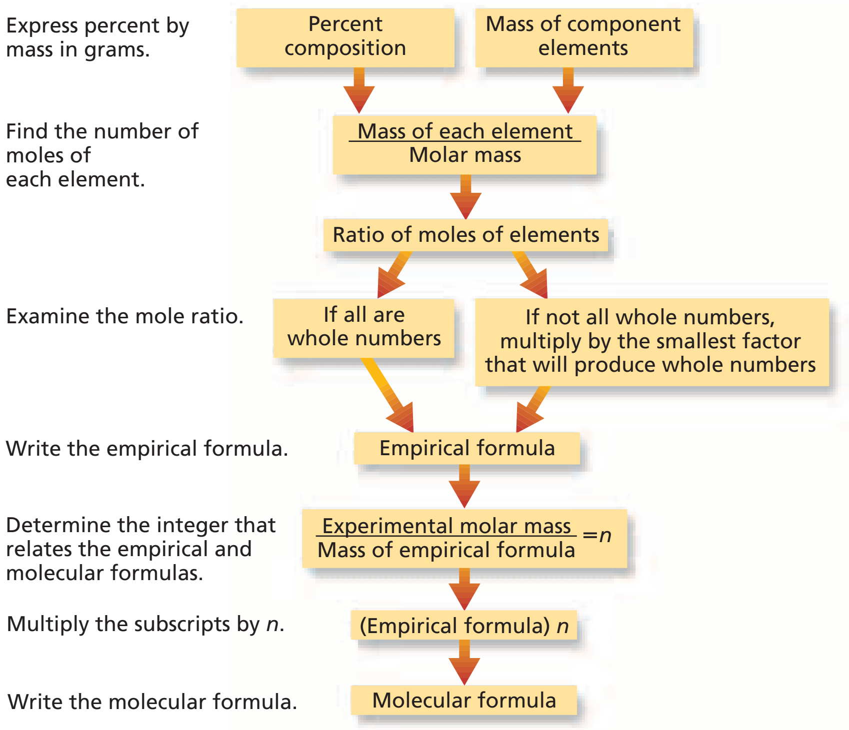
56. The pain reliever morphine contains 17.900 g C, 1.680 g H, 4.225 g O, and 1.228 g N. Determine the empirical formula.

57. An oxide of aluminum contains 0.545 g Al and 0.485 g O. Find the empirical formula for the oxide.

Practice!

For more practice calculating an empirical formula from mass data, go to **Supplemental Practice Problems** in Appendix A.

The steps in determining empirical and molecular formulas from percent composition or mass data are outlined below. As in other calculations, the route leads from mass through moles because formulas are based on the relative numbers of moles of elements in each mole of compound.



Section 11.4 Assessment

58. Explain how percent composition data for a compound are related to the masses of the elements in the compound.

59. What is the difference between an empirical formula and a molecular formula?

60. Explain how you can find the mole ratio in a chemical compound.

61. **Thinking Critically** An analysis for copper was performed on two pure solids. One solid was found to contain 43.0% copper; the other contained 32.0% copper. Could these solids be samples of the same copper-containing compound? Explain your answer.

62. **Inferring** Hematite (Fe_2O_3) and magnetite (Fe_3O_4) are two ores used as sources of iron. Which ore provides the greater percent of iron per kilogram?

Section 11.5 The Formula for a Hydrate

Objectives

- **Explain** what a hydrate is and how its name reflects its composition.
- **Determine** the formula for a hydrate from laboratory data.

Vocabulary

hydrate

Have you ever watched crystals slowly form from a water solution? Sometimes water molecules adhere to the ions as the solid forms. These water molecules become part of the crystal and are called water of hydration. Solids in which water molecules are trapped are called hydrates. A **hydrate** is a compound that has a specific number of water molecules bound to its atoms. **Figure 11-12** shows examples of a common gemstone called opal, which is composed of silicon dioxide (SiO_2). The unusual coloring is the result of the presence of water in the mineral.

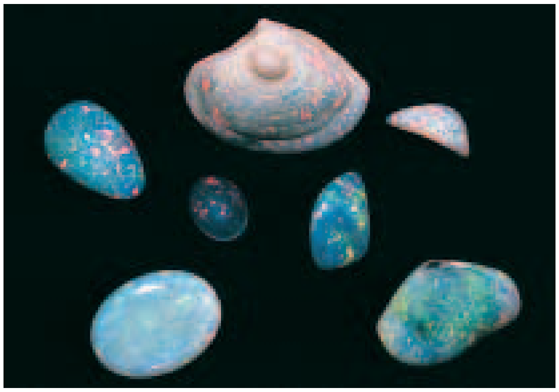


Figure 11-12

The presence of water and various mineral impurities account for the variety of different colored opals. Further changes in color occur when opals are allowed to dry out.

Naming Hydrates

In the formula for a hydrate, the number of water molecules associated with each formula unit of the compound is written following a dot: for example, $Na_2CO_3 \cdot 10H_2O$. This compound is called sodium carbonate decahydrate. In the word *decahydrate*, the prefix *deca-* means ten and the root word *hydrate* refers to water. Decahydrate means that ten molecules of water are associated with one formula unit of compound. The mass of water associated with a formula unit must be included in molar mass calculations. Hydrates are found with a variety of numbers of water molecules. **Table 11-1** lists some common hydrates.

Table 11-1

Formulas for Hydrates

Prefix	Molecules H_2O	Formula	Name
Mono-	1	$(NH_4)_2C_2O_4 \cdot H_2O$	Ammonium oxalate monohydrate
Di-	2	$CaCl_2 \cdot 2H_2O$	Calcium chloride dihydrate
Tri-	3	$NaC_2H_3O_2 \cdot 3H_2O$	Sodium acetate trihydrate
Tetra-	4	$FePO_4 \cdot 4H_2O$	Iron(III) phosphate tetrahydrate
Penta-	5	$CuSO_4 \cdot 5H_2O$	Copper(II) sulfate pentahydrate
Hexa-	6	$CoCl_2 \cdot 6H_2O$	Cobalt(II) chloride hexahydrate
Hepta-	7	$MgSO_4 \cdot 7H_2O$	Magnesium sulfate heptahydrate
Octa-	8	$Ba(OH)_2 \cdot 8H_2O$	Barium hydroxide octahydrate
Deca-	10	$Na_2CO_3 \cdot 10H_2O$	Sodium carbonate decahydrate

Analyzing a Hydrate

To analyze hydrates, you must drive off the water of hydration. Often this is done by heating the compound. The substance remaining after heating is anhydrous, or "without water." For example, hydrated cobalt(II) chloride is a pink solid that turns a deep blue when the water of hydration is driven off and anhydrous cobalt(II) chloride is produced. See **Figure 11-13.**

Figure 11-13

Hydrated $CoCl_2$, shown on the left, is pink. Its anhydrous form, on the right, is blue. The transition from pink to blue was accomplished by heating the hydrate until all water of hydration was removed.

Formula for a hydrate How can you determine the formula for a hydrate? You must find the number of moles of water associated with one mole of the hydrate. Suppose you have a 5.00-g sample of a hydrate of barium chloride. You know that the formula is $BaCl_2 \cdot xH_2O$. You must determine x, the coefficient of H_2O in the hydrate formula that indicates the number of moles of water associated with one mole of $BaCl_2$. To find x, you would heat the sample of the hydrate to drive off the water of hydration. After heating, the dried substance, which is anhydrous $BaCl_2$, has a mass of 4.26 g. The mass of the water of hydration is the difference between the mass of the hydrate (5.00 g) and the mass of the anhydrous compound (4.26 g).

$$5.00 \text{ g } BaCl_2 \text{ hydrate} - 4.26 \text{ g anhydrous } BaCl_2 = 0.74 \text{ g } H_2O$$

You now know the masses of $BaCl_2$ and H_2O in the sample. You can convert these masses to moles using the molar masses. The molar mass of $BaCl_2$ is 208.23 g/mol and the molar mass of H_2O is 18.02 g/mol.

$$4.26 \cancel{\text{g } BaCl_2} \times \frac{1 \text{ mol } BaCl_2}{208.23 \cancel{\text{g } BaCl_2}} = 0.0205 \text{ mol } BaCl_2$$

$$0.74 \cancel{\text{g } H_2O} \times \frac{1 \text{ mol } H_2O}{18.02 \cancel{\text{g } H_2O}} = 0.041 \text{ mol } H_2O$$

Now that the moles of $BaCl_2$ and H_2O have been determined, you can calculate the ratio of moles of H_2O to moles of $BaCl_2$ which is x, the coefficient that precedes H_2O in the formula for the hydrate.

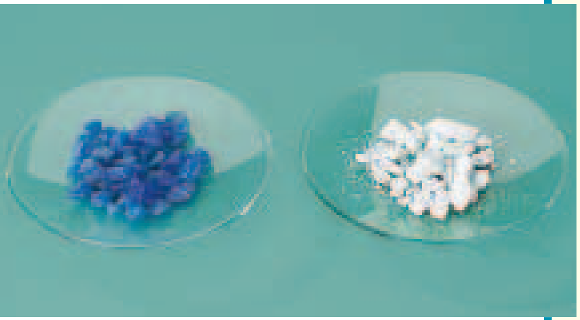
$$x = \frac{\text{moles } H_2O}{\text{moles } BaCl_2} = \frac{0.041 \text{ mol } H_2O}{0.0205 \text{ mol } BaCl_2} = \frac{2.0 \text{ mol } H_2O}{1.00 \text{ mol } BaCl_2} = \frac{2}{1}$$

The ratio of moles of H_2O to moles of $BaCl_2$ is 2:1, so two moles of water are associated with one mole of barium chloride. The value of the coefficient x is 2 and the formula for the hydrate is $BaCl_2 \cdot 2H_2O$. What is the name of the hydrate? The **CHEMLAB** at the end of this chapter will give you experience determining the formula of a hydrate.

EXAMPLE PROBLEM 11-14

Determining the Formula for a Hydrate

A mass of 2.50 g of blue, hydrated copper sulfate ($CuSO_4{\cdot}xH_2O$) is placed in a crucible and heated. After heating, 1.59 g white anhydrous copper sulfate ($CuSO_4$) remains. What is the formula for the hydrate? Name the hydrate.



Heating blue anhydrous copper sulfate drives off the water of hydration and converts it to white anhydrous copper sulfate. How could you convert anhydrous copper sulfate to its blue hydrated form?

1. Analyze the Problem

You are given a mass of hydrated copper sulfate. The mass after heating is the mass of the anhydrous compound. You know the formula for the compound except for *x*, the number of moles of water of hydration.

Known

mass of hydrated compound = 2.50 g $CuSO_4{\cdot}xH_2O$

mass of anhydrous compound = 1.59 g $CuSO_4$

molar mass = 18.02 g/mol H_2O

molar mass = 159.6 g/mol $CuSO_4$

Unknown

formula for hydrate = ?

name of hydrate = ?

2. Solve for the Unknown

Subtract the mass of the anhydrous copper sulfate from the mass of the hydrated copper sulfate to determine the mass of water lost.

mass of hydrated copper sulfate	2.50 g
mass of anhydrous copper sulfate	−1.59 g
mass of water lost	0.91 g

Calculate the number of moles of H_2O and anhydrous $CuSO_4$ using the conversion factor that relates moles and mass based on the molar mass.

$$1.59\ \cancel{\text{g CuSO}_4} \times \frac{1\ \text{mol CuSO}_4}{159.6\ \cancel{\text{g CuSO}_4}} = 0.00996\ \text{mol CuSO}_4$$

$$0.91\ \cancel{\text{g H}_2\text{O}} \times \frac{1\ \text{mol H}_2\text{O}}{18.02\ \cancel{\text{g H}_2\text{O}}} = 0.050\ \text{mol H}_2\text{O}$$

Determine the value of *x*.

$$x = \frac{\text{moles H}_2\text{O}}{\text{moles CuSO}_4}$$

$$x = \frac{0.050\ \text{mol H}_2\text{O}}{0.00996\ \text{mol CuSO}_4} = \frac{5.0\ \text{mol H}_2\text{O}}{1.00\ \text{mol CuSO}_4} = 5$$

The ratio of H_2O to $CuSO_4$ is 5 : 1, so the formula for the hydrate is $CuSO_4{\cdot}5H_2O$, copper(II) sulfate pentahydrate.

3. Evaluate the Answer

Copper(II) sulfate pentahydrate is listed as a hydrate in **Table 11-1**.

PRACTICE PROBLEMS

63. A hydrate is found to have the following percent composition: 48.8% $MgSO_4$ and 51.2% H_2O. What is the formula and name for this hydrate?

64. **Figure 11-13** shows a common hydrate of cobalt(II) chloride. If 11.75 g of this hydrate is heated, 9.25 g of anhydrous cobalt chloride remains. What is the formula and name for this hydrate?

Practice!

For more practice calculating the formula for a hydrate, go to **Supplemental Practice Problems** in Appendix A.

Uses of hydrates The ability of the anhydrous form of a hydrate to absorb water into its crystal structure has some important applications. Anhydrous calcium chloride and calcium sulfate are used as desiccants or drying agents in the laboratory because they can absorb water from the air or from their liquid surroundings. For example, calcium sulfate is often added to solvents such as ethanol and ethyl ether to keep them free of water. Anhydrous calcium chloride is placed in the bottom of tightly sealed containers called desiccators. The calcium chloride absorbs moisture from the air inside the desiccator, thus creating a dry atmosphere in which other substances can be placed to be kept dry. Calcium chloride forms a monohydrate, a dihydrate, and a hexahydrate. Electronic and optical equipment, particularly that transported overseas by ship, is packaged with packets of desiccants that absorb water from the air and prevent moisture from interfering with sensitive circuitry. Some of these uses are illustrated in **Figure 11-14**.

Some hydrates, sodium sulfate decahydrate ($Na_2SO_4 \cdot 10H_2O$) for example, are used to store solar energy. When the Sun's energy heats the hydrate to a temperature greater than 32°C, the single formula unit of Na_2SO_4 in the hydrate dissolves in the 10 moles of water of hydration. In the process, energy is absorbed by the hydrate. This solar energy, stored in the solution of the hydrate, is released when the temperature decreases and the hydrate crystallizes again.

Figure 11-14

Calcium sulfate is not soluble in ethanol so it remains on the bottom of the ethanol bottle and absorbs any water dissolved in the ethanol. Calcium chloride, in the bottom of the desiccator, keeps the air inside the desiccator dry. Porous packets of desiccants can be packaged with materials that need to be kept moisture free.

Section 11.5 Assessment

65. What is a hydrate? What does its name indicate about its composition?

66. Describe the experimental procedure for determining the formula for a hydrate. Explain the reason for each step.

67. Name the compound having the formula $SrCl_2 \cdot 6H_2O$.

68. **Thinking Critically** Explain how the hydrate illustrated in **Figure 11-13** might be used as a means of roughly determining the probability of rain.

69. **Sequencing** Arrange these hydrates in order of increasing percent water content: $MgSO_4 \cdot 7H_2O$, $Ba(OH)_2 \cdot 8H_2O$, $CoCl_2 \cdot 6H_2O$.

Hydrated Crystals

Hydrates are compounds that incorporate water molecules in their crystalline structures. The ratio of moles of water to one mole of the compound is a small whole number. For example, in the hydrated compound copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$), the ratio is 5:1. The ratio of moles of water to one mole of a hydrate can be determined experimentally by heating the hydrate to remove water.

Problem

How can you determine the moles of water in a mole of a hydrated compound?

Objectives

- **Heat** a known mass of hydrated compound until the water is removed.
- **Calculate** the formula for a hydrate using the mass of the hydrated compound and the mass of the anhydrous compound.

Materials

Bunsen burner
ring stand and ring
crucible and lid
clay triangle
crucible tongs
balance
Epsom salts (hydrated $MgSO_4$)
spatula
spark lighter or matches

Safety Precautions

- **Always wear safety goggles and a lab apron.**
- **Hot objects will not appear to be hot.**
- **Use the Bunsen burner carefully.**
- **Turn off the Bunsen burner when not in use.**

Pre-Lab

1. Read the entire **CHEMLAB**.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
3. Explain how you will obtain the mass of water and the mass of anhydrous $MgSO_4$ contained in the hydrate.
4. How will you convert the masses of anhydrous $MgSO_4$ and water to moles?
5. How can you obtain the formula for the hydrate from the moles of anhydrous $MgSO_4$ and the moles of water?

Mass Data and Observations of Epsom Salts	
Observations of hydrated $MgSO_4$	
Mass of crucible and lid	
Mass of crucible, lid, and hydrated $MgSO_4$	
Mass of hydrated $MgSO_4$	
Mass of crucible, lid, and anhydrous $MgSO_4$	
Mass of anhydrous $MgSO_4$	
Mass of water in hydrated $MgSO_4$	
Moles of anhydrous $MgSO_4$	
Moles of water in hydrated $MgSO_4$	
Observation of anhydrous $MgSO_4$	

Procedure

1. Measure to the nearest 0.01 g the mass of a clean, dry crucible with a lid. Record the mass.
2. Add about 3 g hydrated $MgSO_4$ to the crucible. Measure the mass of the crucible, lid, and hydrate to the nearest 0.01 g and record the mass.
3. Record your observations of the hydrate.
4. Place the triangle on the ring of the ring stand. Carefully place the crucible in the triangle as shown in the photo.
5. Place the crucible lid on the crucible slightly cocked to help prevent spattering and allow vapor to escape. Begin heating with a low flame, then gradually progress to a stronger flame. Heat for about 10 minutes.
6. When heating is complete, remove the crucible using tongs. Place the lid on the crucible and allow the crucible and contents to cool.
7. Measure the mass of the crucible, lid, and $MgSO_4$ and record the mass in the data table.
8. Observe the anhydrous $MgSO_4$ and record your observations.

Cleanup and Disposal

1. Discard the anhydrous $MgSO_4$ in a trash container or as directed by your teacher.
2. Return all lab equipment to its proper place and clean your lab station.
3. Wash your hands thoroughly when all lab work and cleanup are complete.

Analyze and Conclude

1. **Using Numbers** Use your experimental data to calculate the formula for hydrated $MgSO_4$.
2. **Observing and Inferring** How did your observations of the hydrated $MgSO_4$ crystals compare with those of the anhydrous $MgSO_4$ crystals?
3. **Drawing Conclusions** Why might the method used in this experiment not be suitable for determining the water of hydration for all hydrates?
4. **Error Analysis** What is the percent error of your calculation of the water of hydration for $MgSO_4$ if the formula for the hydrate is $MgSO_4{\cdot}7H_2O$? What changes would you make in the procedure to reduce error?
5. **Predicting** What might you observe if the anhydrous crystals were left uncovered overnight?

Real-World Chemistry

1. Packets of the anhydrous form of a hydrate are sometimes used to keep cellars from being damp. Is there a limit to how long a packet could be used?
2. Gypsum ($CaSO_4{\cdot}2H_2O$) is a mineral used for making wallboard for construction. The mineral is stripped of three-quarters of its water of hydration in a process called calcinning. Then, after mixing with water, it hardens to a white substance called plaster of paris. Infer what happens as calcinned gypsum becomes plaster of paris.

CHEMISTRY and Technology

Making Medicines

In 1859, a chemist named Herman Kolbe identified the chemical structure of the compound that made willow bark an effective herbal medicine. Unfortunately, many people who took the medicine suffered upset stomachs from it. One such person was the father of a chemist named Felix Hoffman. To help his father, Hoffman changed the compound in such a way that it eased his father's arthritis without irritating his stomach. The new compound was acetylsalicylic acid, now known as aspirin.

Chemical Synthesis

Like aspirin, many medicines are compounds synthesized in the laboratory. During chemical synthesis, chemists combine simple compounds to form more complex compounds. Often, several chemical reactions must occur in a particular sequence to produce the desired product. The process is like putting building blocks together one at a time as the chemist carries out several chemical reactions in a particular sequence. The final product is then tested, or screened, to see if it has the desired effect. In the case of a medicine, the compound must be designed to interact with a molecule in the body known as the target molecule. The compound must have a particular shape and characteristics to be effective against the target molecule. If chemists know the shape and chemical properties of the target molecule, they try to design a compound to fit it exactly. More often, chemists must try as many compounds as possible until they find one that works.

Combinatorial Chemistry

In the past, finding a potentially successful medicine would take many years. Now, chemists may be able to speed up the process by turning to the techniques of combinatorial chemistry. In combinatorial chemistry, chemists use robots to put chemical building blocks together in every possible combination at the same time. Instead of chemists making one compound per week, combinatorial chemistry enables chemists to produce as many as 100 compounds a day and up to tens of thousands of compounds a year!

Looking to the Future

The goal of combinatorial chemistry is not to produce as many compounds as possible but to try to identify a basic structure that is effective and then develop variations of that structure. It's too soon to tell if combinatorial chemistry will revolutionize the way medicines are developed because the process has been used only within the last decade. However, some promising leads have been generated in months rather than years, and compounds produced from combinatorial chemistry are currently being tested.

Investigating the Technology

1. **Controlling Variables** Why would it be important for chemists to control the variables of the reactions involved in combinatorial chemistry?

2. **Using Resources** Research the two main approaches to combinatorial chemistry. Find out how they are alike and how they are different.

ChemistryOnline

Visit the Chemistry Web site at **chemistrymc.com** to find links to more information about combinatorial chemistry.

Summary

11.1 Measuring Matter

- The mole is a unit used to count particles indirectly. One mole is the amount of a pure substance that contains 6.02×10^{23} representative particles.
- One mole of carbon-12 atoms contains 12 grams of the isotope carbon-12.

11.2 Mass and the Mole

- The molar mass of an element is the numerical equivalent of the atomic mass (amu) in grams.
- The molar mass of any substance is the mass in grams of Avogadro's number of representative particles of the substance.
- Molar mass is used to convert from moles of an element to mass, and the inverse of molar mass is used to convert from mass of an element to moles.

11.3 Moles of Compounds

- Subscripts in a chemical formula indicate how many moles of each element are in one mole of the compound.
- The molar mass of a compound is the sum of the masses of all the moles of elements present in the compound.

11.4 Empirical and Molecular Formulas

- The percent composition of a known compound can be calculated by dividing the mass of each element in one mole by the mass of a mole of the compound and multiplying by 100.
- The subscripts in an empirical formula are in a ratio of the smallest whole numbers of moles of the elements in the compound.
- The molecular formula for a compound can be determined by finding the integer by which the mass of the empirical formula differs from the molar mass of the compound.

11.5 The Formula for a Hydrate

- The formula for a hydrate consists of the formula for the ionic compound and the number of water molecules associated with one formula unit.
- The name of a hydrate consists of the compound name followed by the word *hydrate* with a prefix indicating the number of water molecules associated with one mole of compound.
- Anhydrous compounds are formed when hydrates are heated and the water of hydration is driven off.

Key Equations and Relationships

- $\text{number of representative particles} = \text{number of moles} \times \dfrac{6.02 \times 10^{23}\ \text{representative particles}}{1\ \text{mole}}$
  (p. 311)
- $\text{number of moles} = \text{number of representative particles} \times \dfrac{1\ \text{mole}}{6.02 \times 10^{23}\ \text{representative particles}}$
  (p. 311)
- $\text{mass} = \text{number of moles} \times \dfrac{\text{number of grams}}{1\ \text{mole}}$
  (p. 315)
- $\text{number of moles} = \text{mass} \times \dfrac{1\ \text{mole}}{\text{number of grams}}$
  (p. 316 Example Problem)
- $\text{percent by mass} = \dfrac{\text{mass of element}}{\text{mass of compound}} \times 100$
  (p. 328)
- $\text{molecular formula} = (\text{empirical formula})n$
  (p. 334)

Vocabulary

- Avogadro's number (p. 310)
- empirical formula (p. 331)
- hydrate (p. 338)
- molar mass (p. 313)
- mole (p. 310)
- molecular formula (p. 333)
- percent composition (p. 328)

chemistrymc.com/vocabulary_puzzlemaker

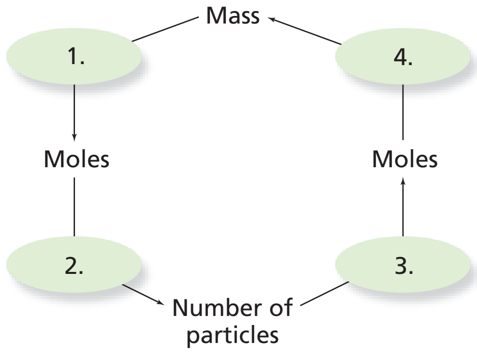
CHAPTER 11 ASSESSMENT

Chemistry Online

Go to the Chemistry Web site at **chemistrymc.com** *for additional Chapter 11 Assessment.*

Concept Mapping

70. Complete this concept map by placing in each box the conversion factor needed to convert from each measure of matter to the next.



Mastering Concepts

71. Why is the mole an important unit to chemists? (11.1)

72. How is a mole similar to a dozen? (11.1)

73. What is the numerical value of Avogadro's number? (11.1)

74. What is molar mass? (11.2)

75. Which contains more atoms, a mole of silver atoms or a mole of gold atoms? Explain your answer. (11.2)

76. Discuss the relationships that exist between the mole, molar mass, and Avogadro's number. (11.2)

77. Which has a greater mass, a mole of silver atoms or a mole of gold atoms? Explain your answer. (11.2)

78. Explain the difference between atomic mass (amu) and molar mass (gram). (11.2)

79. If you divide the molar mass of an element by Avogadro's number, what is the meaning of the quotient? (11.2)

80. List three conversion factors used in molar conversions. (11.3)

81. What information is provided by the formula for potassium chromate (K_2CrO_4)? (11.3)

82. Which of the following molecules contains the most moles of carbon atoms per mole of the compound: ascorbic acid ($C_6H_8O_6$), glycerin ($C_3H_8O_3$), or vanillin ($C_8H_8O_3$)? Explain. (11.3)

83. Explain what is meant by percent composition. (11.4)

84. What is the difference between an empirical formula and a molecular formula? Use an example to illustrate your answer. (11.4)

85. Do all pure samples of a given compound have the same percent composition? Explain. (11.4)

86. What information must a chemist obtain in order to determine the empirical formula of an unknown compound? (11.4)

87. What is a hydrated compound? Use an example to illustrate your answer. (11.5)

88. Explain how hydrates are named. (11.5)

Mastering Problems

Mole-Particle Conversions (11.1)

89. Determine the number of representative particles in each of the following.

a. 0.250 mol silver
b. 8.56×10^{-3} mol sodium chloride
c. 35.3 mol carbon dioxide
d. 0.425 mol nitrogen (N_2)

90. Determine the number of moles in each of the following.

a. 3.25×10^{20} atoms lead
b. 4.96×10^{24} molecules glucose
c. 1.56×10^{23} formula units sodium hydroxide
d. 1.25×10^{25} copper(II) ions

91. Make the following conversions.

a. 1.51×10^{15} atoms Si to mol Si
b. 4.25×10^{-2} mol H_2SO_4 to molecules H_2SO_4
c. 8.95×10^{25} molecules CCl_4 to mol CCl_4
d. 5.90 mol Ca to atoms Ca

92. How many molecules are contained in each of the following?

a. 1.35 mol carbon disulfide (CS_2)
b. 0.254 mol diarsenic trioxide (As_2O_3)
c. 1.25 mol water
d. 150.0 mol HCl

93. How many moles contain each of the following?

a. 1.25×10^{15} molecules carbon dioxide
b. 3.59×10^{21} formula units sodium nitrate
c. 2.89×10^{27} formula units calcium carbonate

94. A bracelet containing 0.200 mol of metal atoms is 75% gold. How many particles of gold atoms are in the bracelet?

95. A solution containing 0.250 mol Cu^{2+} ions is added to another solution containing 0.130 mol Ca^{2+} ions. What is the total number of metal ions in the combined solution?

96. If a snowflake contains 1.9×10^{18} molecules of water, how many moles of water does it contain?

97. If you could count two atoms every second, how long would it take you to count a mole of atoms? Assume that you counted continually 24 hours every day. How does the time you calculated compare with the age of Earth, which is estimated to be 4.5×10^{9} years old?

Mole-Mass Conversions (11.2)

98. Calculate the mass of the following.

a. 5.22 mol He
b. 0.0455 mol Ni
c. 2.22 mol Ti
d. 0.00566 mol Ge

99. Make the following conversions.

a. 3.50 mol Li to g Li
b. 7.65 g Co to mol Co
c. 5.62 g Kr to mol Kr
d. 0.0550 mol As to g As

100. Determine the mass in grams of the following.

a. 1.33×10^{22} mol Sb
b. 4.75×10^{14} mol Pt
c. 1.22×10^{23} mol Ag
d. 9.85×10^{24} mol Cr

Particle-Mass Conversions (11.2)

101. Convert the following to mass in grams.

a. 4.22×10^{15} atoms U
b. 8.65×10^{25} atoms H
c. 1.25×10^{22} atoms O
d. 4.44×10^{23} atoms Pb

102. A sensitive balance can detect masses of 1×10^{-8} g. How many atoms of silver would be in a sample having this mass?

103. Calculate the number of atoms in each of the following.

a. 25.8 g Hg
b. 0.0340 g Zn
c. 150 g Ar
d. 0.124 g Mg

104. Which has more atoms, 10.0 g of carbon or 10.0 g of calcium? How many atoms does each have?

105. Which has more atoms, 10.0 moles of carbon or 10.0 moles of calcium? How many does each have?

106. A mixture contains 0.250 mol Fe and 1.20 g C. What is the total number of atoms in the mixture?

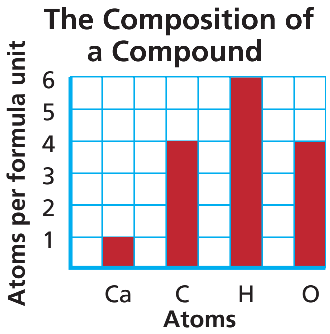
Chemical Formulas (11.3)

107. In the formula for sodium phosphate (Na_3PO_4) how many moles of sodium are represented? How many moles of phosphorus? How many moles of oxygen?

108. How many moles of oxygen atoms are contained in the following?

a. 2.50 mol $KMnO_4$
b. 45.9 mol CO_2
c. 1.25×10^{-2} mol $CuSO_4 \cdot 5H_2O$

109. The graph shows the numbers of atoms of each element in a compound. What is its formula? What is its molar mass?



110. How many carbon tetrachloride molecules are in 3.00 mol carbon tetrachloride (CCl_4)? How many carbon atoms? How many chlorine atoms? How many total atoms?

Molar Mass (11.3)

111. Determine the molar mass of each of the following.

a. nitric acid (HNO_3)
b. ammonium nitrate (NH_4NO_3)
c. zinc oxide (ZnO)
d. cobalt chloride ($CoCl_2$)

112. Calculate the molar mass of each of the following.

a. ascorbic acid ($C_6H_8O_6$)
b. sulfuric acid (H_2SO_4)
c. silver nitrate ($AgNO_3$)
d. saccharin ($C_7H_5NO_3S$)

113. Determine the molar mass of allyl sulfide, the compound responsible for the smell of garlic. The chemical formula of allyl sulfide is $(C_3H_5)_2S$.

Mass-Mole Conversions (11.3)

114. How many moles are in 100.0 g of each of the following compounds?

a. dinitrogen oxide (N_2O)
b. methanol (CH_3OH)

115. What is the mass of each of the following?

a. 4.50×10^{-2} mol $CuCl_2$
b. 1.25×10^{2} mol $Ca(OH)_2$

116. Determine the number of moles in each of the following.

a. 1.25×10^{2} g Na_2S **b.** 0.145 g H_2S

117. Benzoyl peroxide is a substance used as an acne medicine. What is the mass in grams of 3.50×10^{-2} moles of benzoyl peroxide ($C_{14}H_{10}O_4$)?

118. Hydrofluoric acid is a substance used to etch glass. Determine the mass of 4.95×10^{25} HF molecules.

119. How many moles of aluminum ions are in 45.0 g of aluminum oxide?

120. How many moles of ions are in the following?

a. 0.0200 g $AgNO_3$ **c.** 0.500 g $Ba(OH)_2$
b. 0.100 mol K_2CrO_4 **d.** 1.00×10^{-9} mol Na_2CO_3

Mass-Particle Conversions (11.3)

121. Calculate the values that will complete the table.

Table 11-2

Moles, Mass, and Representative Particles			
Compound	**Number of moles**	**Mass (g)**	**Representative particles**
Silver acetate $Ag(C_2H_3O_2)$	2.50		
Glucose $C_6H_{12}O_6$		324.0	
Benzene C_6H_6			5.65×10^{21}
Lead(II) sulfide PbS		100.0	

122. How many formula units are present in 500.0 g lead(II) chloride?

123. Determine the number of atoms in 3.50 g gold.

124. Calculate the mass of 3.62×10^{24} molecules of glucose ($C_6H_{12}O_6$).

125. Determine the number of molecules of ethanol (C_2H_5OH) in 47.0 g.

126. What mass of iron(III) chloride contains 2.35×10^{23} chloride ions?

127. How many moles of iron can be recovered from 100.0 kg Fe_3O_4?

128. The mass of an electron is 9.11×10^{-28} g. What is the mass of a mole of electrons?

129. Vinegar is 5.0% acetic acid (CH_3COOH). How many molecules of acetic acid are present in 25.0 g vinegar?

130. The density of lead is 11.3 g/cm^3. Calculate the volume of one mole lead.

131. Calculate the moles of aluminum ions present in 250.0 g aluminum oxide (Al_2O_3).

132. Determine the number of chloride ions in 10.75 g of magnesium chloride.

133. Acetaminophen, a common aspirin substitute, has the formula $C_8H_9NO_2$. Determine the number of molecules of acetaminophen in a 500 mg tablet.

134. Calculate the number of sodium ions present in 25.0 g sodium chloride.

135. Determine the number of oxygen atoms present in 25.0 g carbon dioxide.

Percent Composition (11.4)

136. Express the composition of each of the following as the mass percent of its elements (percent composition).

a. sucrose ($C_{12}H_{22}O_{11}$)
b. magnetite (Fe_3O_4)
c. aluminum sulfate ($Al_2(SO_4)_3$)

137. Which of the following iron compounds contain the greatest percentage of iron: pyrite (FeS_2), hematite (Fe_2O_3), or siderite ($FeCO_3$)?

138. Determine the empirical formula for each of the following compounds.

a. ethylene (C_2H_4)
b. ascorbic acid ($C_6H_8O_6$)
c. naphthalene ($C_{10}H_8$)

139. Caffeine, a stimulant found in coffee, has the chemical formula $C_8H_{10}N_4O_2$.

a. Calculate the molar mass of caffeine.
b. Determine the percent composition of caffeine.

140. Which of the titanium-containing minerals rutile (TiO_2) or ilmenite ($FeTiO_3$) has the larger percent of titanium?

141. Vitamin E, found in many plants, is thought to retard the aging process in humans. The formula for vitamin E is $C_{29}H_{50}O_2$. What is the percent composition of vitamin E?

142. Aspartame, an artificial sweetener, has the formula $C_{14}H_{18}N_2O_5$. Determine the percent composition of aspartame.

Empirical and Molecular Formulas (11.4)

143. The hydrocarbon used in the manufacture of foam plastics is called styrene. Analysis of styrene indicates the compound is 92.25% C and 7.75% H and has a molar mass of 104 g/mol. Determine the molecular formula for styrene.

144. Monosodium glutamate (MSG) is sometimes added to food to enhance flavor. Analysis determined this compound to be 35.5% C, 4.77% H, 8.29% N, 13.6% Na, and 37.9% O. What is the empirical formula for MSG?

145. Determine the molecular formula for ibuprofen, a common headache remedy. Analysis of ibuprofen yields a molar mass of 206 g/mol and a percent composition of 75.7% C, 8.80% H and 15.5% O.

146. Vanadium oxide is used as an industrial catalyst. The percent composition of this oxide is 56.0% vanadium and 44.0% oxygen. Determine the empirical formula for vanadium oxide.

147. What is the empirical formula of a compound that contains 10.52 g Ni, 4.38 g C, and 5.10 g N?

148. The Statue of Liberty turns green in air because of the formation of two copper compounds, $Cu_3(OH)_4SO_4$ and $Cu_4(OH)_6SO_4$. Determine the mass percent of copper in these compounds.

149. Analysis of a compound containing chlorine and lead reveals that the compound is 59.37% lead. The molar mass of the compound is 349.0 g/mol. What is the empirical formula for the chloride? What is the molecular formula?

150. Glycerol is a thick, sweet liquid obtained as a byproduct of the manufacture of soap. Its percent composition is 39.12% carbon, 8.75% hydrogen, and 52.12% oxygen. The molar mass is 92.11 g/mol. What is the molecular formula for glycerol?

The Formula for a Hydrate (11.5)

151. Determine the mass percent of anhydrous sodium carbonate (Na_2CO_3) and water in sodium carbonate decahydrate ($Na_2CO_3 \cdot 10H_2O$).

152. What is the formula and name of a hydrate that is 85.3% barium chloride and 14.7% water?

153. Gypsum is hydrated calcium sulfate. A 4.89-g sample of this hydrate was heated, and after the water was driven off, 3.87 g anhydrous calcium sulfate remained. Determine the formula of this hydrate and name the compound.

154. The table shows data from an experiment to determine the formulas of hydrated barium chloride. Determine the formula for the hydrate and its name.

Table 11-3

Data for $BaCl_2 \cdot xH_2O$	
Mass of empty crucible	21.30 g
Mass of hydrate + crucible	31.35 g
Initial mass of hydrate	
Mass after heating 5 min	29.87 g
Mass of anhydrous solid	

155. A 1.628-g sample of a hydrate of magnesium iodide is heated until its mass is reduced to 1.072 g and all water has been removed. What is the formula of the hydrate?

156. Hydrated sodium tetraborate ($Na_2B_4O_7 \cdot xH_2O$) is commonly called borax. Chemical analysis indicates that this hydrate is 52.8% sodium tetraborate and 47.2% water. Determine the formula and name the hydrate.

Mixed Review

Sharpen your problem-solving skills by answering the following.

157. Determine the following:

a. the number of representative particles in 3.75 g Zn
b. the mass of 4.32×10^{22} atoms Ag
c. the number of sodium ions in 25.0 g of Na_2O

158. Which of the following has the greatest number of oxygen atoms?

a. 17.63 g CO_2
b. 3.21×10^{22} molecules CH_3OH
c. 0.250 mol $C_6H_{12}O_6$

159. Which of the following compounds has the greatest percent of oxygen by mass: TiO_2, Fe_2O_3, Al_2O_3?

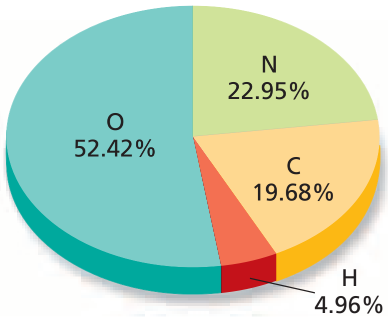
160. Naphthalene, commonly known as moth balls, is composed of 93.7% carbon and 6.3% hydrogen. The molar mass of napthalene is 128 g/mol. Determine the empirical and molecular formulas for naphthalene.

161. Which of the following molecular formulas are also empirical formulas: ethyl ether ($C_4H_{10}O$), aspirin ($C_9H_8O_4$), butyl dichloride ($C_4H_8Cl_2$), glucose ($C_6H_{12}O_6$).

162. Calculate each of the following:

a. the number of moles in 15.5 g Na_2SO_4
b. the number of formula units in 0.255 mol NaCl
c. the mass in grams of 0.775 mol SF_6
d. the number of Cl^- ions in 14.5 g $MgCl_2$.

163. The graph shows the percent composition of a compound containing carbon, hydrogen, oxygen, and nitrogen. How many grams of each element are present in 100 g of the compound?



164. A party balloon was filled with 9.80×10^{22} atoms of helium. After 24 hours, 45% of the helium had escaped. How many atoms of helium remained?

165. Tetrafluoroethylene, which is used in the production of Teflon, is composed of 24.0% carbon and 76.0% fluorine and has a molar mass of 100.0 g/mol. Determine the empirical and molecular formulas of this compound.

166. Calculate the mass in grams of one atom of lead.

167. Diamond is a naturally occurring form of carbon. If you have a 0.25-carat diamond, how many carbon atoms are present? (1 carat = 0.200 g)

168. How many molecules of isooctane (C_8H_{18}) are present in 1.00 L? (density of isooctane = 0.680 g/mL)

169. Calculate the number of molecules of water in a swimming pool which is 40.0 m in length, 20.0 m in width, and 5.00 m in depth. Assume that the density of water is 1.00 g/cm^3.

Thinking Critically

170. Analyze and Conclude A mining company has two possible sources of copper: chalcopyrite ($CuFeS_2$) and chalcocite (Cu_2S). If the mining conditions and the extraction of copper from the ore were identical for each of the ores, which ore would yield the greater quantity of copper? Explain your answer.

171. Designing an Experiment Design an experiment that can be used to determine the amount of water in alum ($KAl(SO_4)_2 \cdot xH_2O$).

172. Concept Mapping Design a concept map that illustrates the mole concept. Include moles, Avogadro's number, molar mass, number of particles, percent composition, empirical formula, and molecular formula.

173. Communicating If you use the expression "a mole of nitrogen," is it perfectly clear what you mean, or is there more than one way to interpret the expression? Explain. How could you change the expression to make it more precise?

Writing in Chemistry

174. Octane ratings are used to identify certain grades of gasoline. Research *octane rating* and prepare a pamphlet for consumers identifying the different types of gasoline, the advantages of each, and when each grade is used.

175. Research the life of the Italian chemist Amedeo Avogadro (1776–1856) and how his work led scientists to the number of particles in a mole.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

176. Express the following answers with the correct number of significant figures. (Chapter 2)

a. 18.23 − 456.7
b. 4.233 ÷ 0.0131
c. (82.44 × 4.92) + 0.125

177. Distinguish between atomic number and mass number. How do these two numbers compare for isotopes of an element? (Chapter 4)

178. Write balanced equations for the following reactions. (Chapter 10)

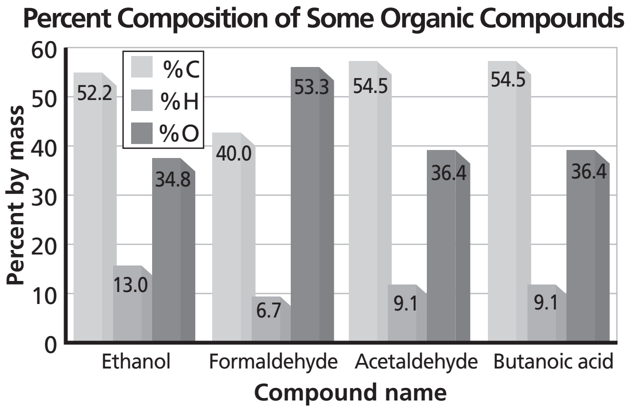
a. Magnesium metal and water combine to form solid magnesium hydroxide and hydrogen gas.
b. Dinitrogen tetroxide gas decomposes into nitrogen dioxide gas.
c. Aqueous solutions of sulfuric acid and potassium hydroxide undergo a double replacement reaction.

179. How can you tell if a chemical equation is balanced? (Chapter 10)

STANDARDIZED TEST PRACTICE
CHAPTER 11

Use these questions and the test-taking tip to prepare for your standardized test.

Interpreting Graphs Use the graph to answer questions 1–4.



1. Acetaldehyde and butanoic acid must have the same

a. molecular formula.
b. empirical formula.
c. molar mass.
d. chemical properties.

2. If the molar mass of butanoic acid is 88.1 g/mol, then what is its molecular formula?

a. CH_2O
b. C_2H_4O
c. C_6HO_2
d. $C_4H_8O_2$

3. What is the empirical formula of ethanol?

a. C_4HO_3
b. $C_{52}H_{13}O_{35}$
c. C_2H_6O
d. $C_4H_{13}O_2$

4. The empirical formula of formaldehyde is the same as its molecular formula. How many grams are in 2.000 moles of formaldehyde?

a. 30.00 g
b. 60.06 g
c. 182.0 g
d. 200.0 g

5. A mole is all of the following EXCEPT

a. the atomic or molar mass of an element or compound.
b. Avogadro's number of molecules of a compound.
c. the number of atoms in exactly 12 g of pure ^{12}C.
d. the SI measurement unit for the amount of a substance.

6. How many atoms are in 0.625 moles of Ge (atomic mass = 72.59 amu)?

a. 2.73×10^{25}
b. 6.99×10^{25}
c. 3.76×10^{23}
d. 9.63×10^{23}

7. What is the molar mass of fluorapatite ($Ca_5(PO_4)_3F$)?

a. 314 g/mol
b. 344 g/mol
c. 442 g/mol
d. 504 g/mol

8. How many moles of cobalt(III) titanate (Co_2TiO_4) are in 7.13 g of the compound?

a. 2.39×10^{1} mol
b. 3.14×10^{-2} mol
c. 3.22×10^{1} mol
d. 4.17×10^{-2} mol

9. Magnesium sulfate ($MgSO_4$) is often added to water-insoluble liquid products of chemical reactions to remove any unwanted water. $MgSO_4$ readily absorbs water to form two different hydrates. One of these hydrates is found to contain 13.0% H_2O and 87.0% $MgSO_4$. What is the name of this hydrate?

a. magnesium sulfate monohydrate
b. magnesium sulfate dihydrate
c. magnesium sulfate hexahydrate
d. magnesium sulfate heptahydrate

10. What is the mass of one molecule of barium hexafluorosilicate ($BaSiF_6$)?

a. 1.68×10^{26} g
b. 2.16×10^{21} g
c. 4.64×10^{-22} g
d. 6.02×10^{-23} g

TEST-TAKING TIP

Your Answers Are Better Than The Test's When you know how to answer a question, answer it in your own words before looking at the answer choices. Often more than one answer choice will look good, so do the calculations first, and arm yourself with your answer before looking. If you need to, cover the choices with one hand until you're ready.