CHAPTER 10 The Mole

BIG Idea The mole represents a large number of extremely small particles.

10.1 Measuring Matter

MAIN Idea Chemists use the mole to count atoms, molecules, ions, and formula units.

10.2 Mass and the Mole

MAIN Idea A mole always contains the same number of particles; however, moles of different substances have different masses.

10.3 Moles of Compounds

MAIN Idea The molar mass of a compound can be calculated from its chemical formula and can be used to convert from mass to moles of that compound.

10.4 Empirical and Molecular Formulas

MAIN Idea A molecular formula of a compound is a whole-number multiple of its empirical formula.

10.5 Formulas of Hydrates

MAIN Idea Hydrates are solid ionic compounds in which water molecules are trapped.

ChemFacts

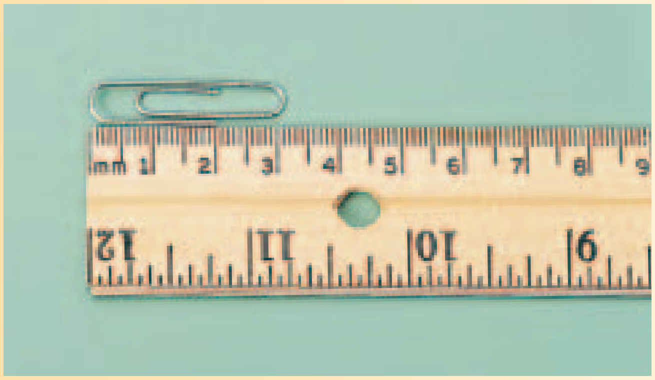
- The U.S. Mint has never officially produced a coin called the "penny"; the official name is the United States one-cent coin.
- The present-day penny is copper-plated zinc, and has a composition of 97.5% Zn and 2.5% Cu.
- The Denver and Philadelphia Mints produce 65 million to 80 million coins a day.

Start-Up Activities

LAUNCH Lab

How much is a mole?

Counting large numbers of items is easier when you use counting units such as decades or dozens. Chemists use a counting unit called the mole.



Procedure

1. Read and complete the lab safety form.
2. Select an item to measure, such as a **paper clip, gum drop,** or **marshmallow,** from the choices provided by your teacher.
   WARNING: ***Do not eat or taste any items used in the lab.***
3. Use a **ruler** to measure the length of your item to the nearest 0.1 cm.

Analysis

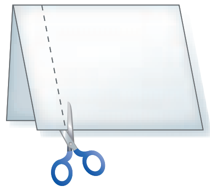
1. **Calculate** If a mole is 6.02×10^{23} items, how far will a mole of your items, placed end-to-end lengthwise, extend into space? Express your answer in meters.
2. **Calculate** Convert the distance in Question 1 to light-years (ly). (1 ly $= 9.46 \times 10^{15}$ m)
3. **Compare** the distance you calculated in Question 2 with these astronomical distances:
   a. distance to nearest star (other than the Sun) = 4.3 ly
   b. distance to the center of our galaxy = 30,000 ly
   c. distance to nearest galaxy $= 2 \times 10^{6}$ ly

Inquiry Compare your item to another used by one of your classmates. Would a mole of your item have the same mass as a mole of the other item? Design an investigation to determine if there is a relationship between mass and moles.

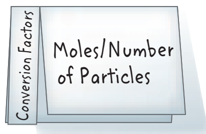
FOLDABLES™ Study Organizer

Conversion Factors Make the following Foldable to help you organize information about conversion factors.

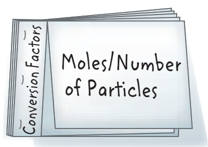
STEP 1 Collect three sheets of paper. Fold each sheet in half. Measure and draw a line about 3 cm from the left edge. Cut along the line to the fold. Repeat for each sheet of paper.



STEP 2 Label each top sheet with a description of the conversion factor.



STEP 3 Staple the sheets together along the outer edge of the narrow flaps.



FOLDABLES **Use this Foldable with Sections 10.1, 10.2, and 10.3.** As you read the sections, record information about conversion factors and summarize the steps involved in each conversion.

Chemistry Online

Visit glencoe.com to:
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Matt Meadows

Section 10.1

Objectives

- **Explain** how a mole is used to indirectly count the number of particles of matter.
- **Relate** the mole to a common everyday counting unit.
- **Convert** between moles and number of representative particles.

Review Vocabulary

molecule: two or more atoms that covalently bond together to form a unit

New Vocabulary

mole
Avogadro's number

Measuring Matter

MAIN Idea Chemists use the mole to count atoms, molecules, ions, and formula units.

Real-World Reading Link Has your class ever had a contest to guess how many pennies or jelly beans were in a jar? You might have noticed that the smaller the object is, the harder it is to count.

Counting Particles

If you were buying a bouquet of roses for a special occasion, you probably would not ask for 12 or 24; you would 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 10.1**—a pair, a dozen, a gross, and a ream—represents a specific number of items. These units make counting objects easier. It is easier to buy and sell paper by the ream—500 sheets—than by the individual sheet.

Each of the counting units shown in **Figure 10.1** is appropriate for certain kinds of objects, depending primarily on their size and function. But regardless of the object—gloves, eggs, pencils, or paper—the number that the unit represents is always constant. Chemists also need a convenient method for accurately counting the number of atoms, molecules, or formula units in a sample of a substance. However, atoms are so small and there are so many of them in even the smallest sample that it is impossible to count them directly. Because of this, chemists created a counting unit called the mole. In the Launch Lab, you probably found that a mole of any object is an enormous number of items.

Figure 10.1 Different units are used to count different types of objects. A pair is two objects, a dozen is 12, a gross is 144, and a ream is 500.
List *What other counting units are you familiar with?*

The mole The **mole,** abbreviated mol, is the SI base unit used to measure the amount of a substance. A mole is defined as the number of carbon atoms in exactly 12 g of pure carbon-12. Through years of experimentation, it has been established that a mole of anything contains 6.0221367×10^{23} representative particles. A representative particle is any kind of particle, such as an atom, a molecule, a formula unit, an electron, or an ion. If you write out Avogadro's number, it looks like this.

$$602{,}213{,}670{,}000{,}000{,}000{,}000{,}000$$

The number 6.0221367×10^{23} is called **Avogadro's number,** in honor of the Italian physicist and lawyer Amedeo Avogadro, who, in 1811, determined the volume of 1 mol of a gas. In this book, Avogadro's number is rounded to three significant figures, 6.02×10^{23}.

To count extremely small particles, such as atoms, Avogadro's number must be an enormous quantity. As you might 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! **Figure 10.2,** however, shows that it is convenient to use the mole to measure amounts of substances. One-mole quantities of water, copper, and salt 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 NaCl formula unit.

VOCABULARY

SCIENCE USAGE V. COMMON USAGE

Mole

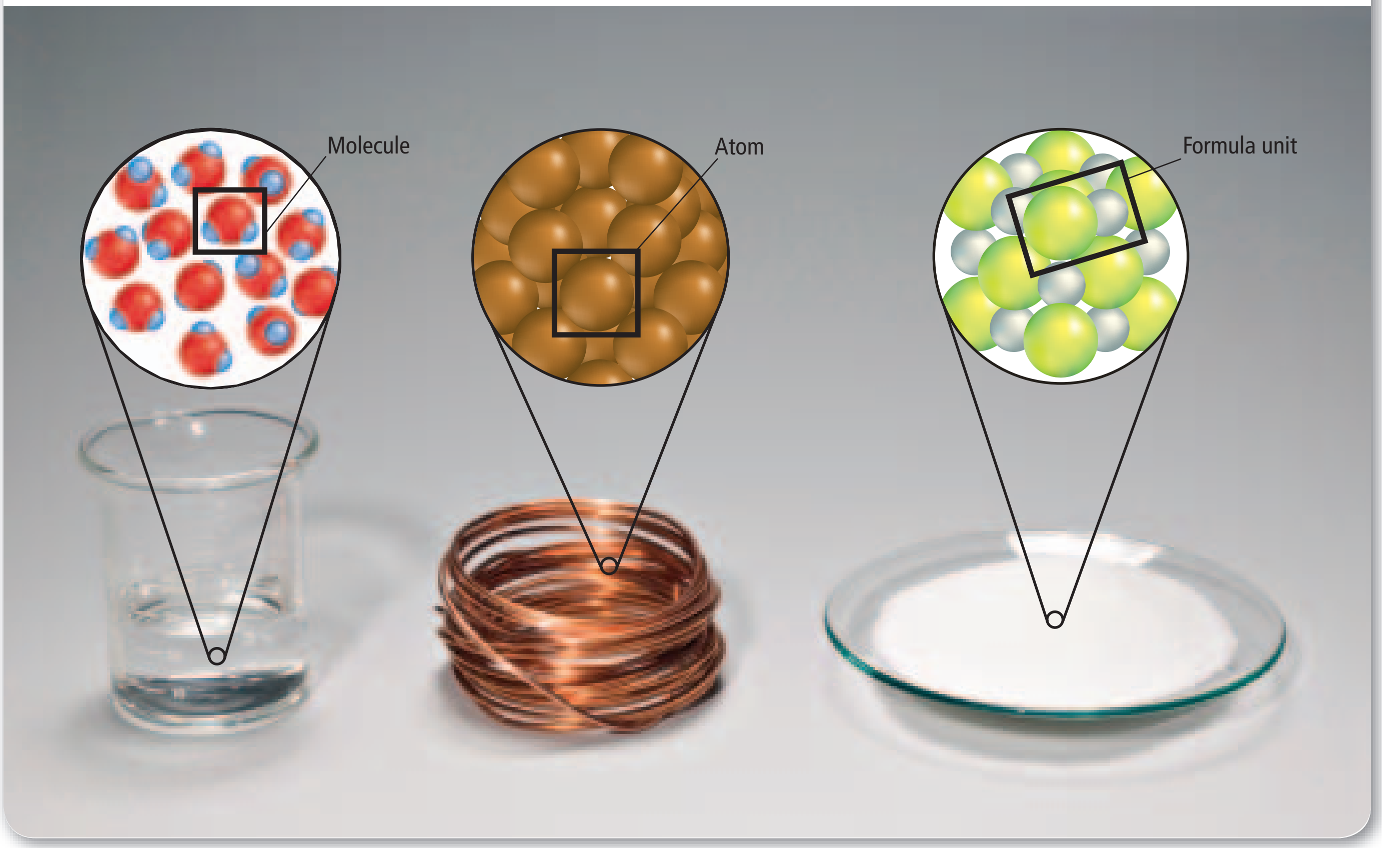
Science usage: an SI base unit used to measure the quantity of matter

The chemist measured out a mole of the compound.

Common usage: a small burrowing animal

The damage to the lawn was caused by a mole.

■ **Figure 10.2** The amount of each substance shown is 6.02×10^{23} or 1 mol of representative particles. The representative particle for each substance is shown in a box. Refer to **Table R-1** on page 968 for a key to atom color conventions.



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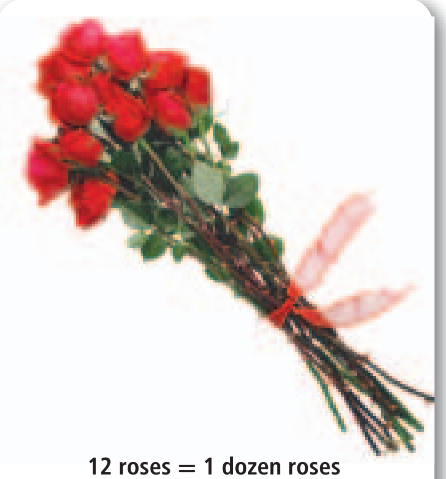


12 roses = 1 dozen roses

Figure 10.3 A key to using dimensional analysis is correctly identifying the mathematical relationship between the units you are converting. The relationship shown here, 12 roses = 1 dozen roses, can be used to write two conversion factors.

Converting Between Moles and Particles

Suppose you buy three-and-one-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). First, identify the mathematical relationship that relates the given unit with the desired unit. **Figure 10.3** shows the relationship.

$$\text{Relationship: } 1 \text{ dozen roses} = 12 \text{ roses}$$

By dividing each side of the equality by the other side, you can write two conversion factors from the relationship.

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

Then choose the conversion factor that, when multiplied by the known quantity, results in the desired unit. When set up correctly, all units cancel except those required for the answer.

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

Here, dozens of roses cancels, leaving roses as the desired unit.

Reading Check **Describe** how you can tell if the wrong conversion factor has been used.

FOLDABLES
Incorporate information from this section into your Foldable.

Moles to particles Now suppose you want to determine how many particles of sucrose are in 3.50 mol of sucrose. The relationship between moles and representative particles is given by Avogadro's number.

$$1 \text{ mol of representative particles} = 6.02 \times 10^{23} \text{ representative particles}$$

Using this relationship, you can write two different conversion factors that relate representative particles and moles.

$$\frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mol}}$$

$$\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ representative particles}}$$

By using the correct conversion factor, you can find the number of representative particles in a given number of moles.

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

As shown in **Figure 10.4,** the representative particle of sucrose is a molecule. To obtain the number of sucrose molecules contained in 3.50 mol of sucrose, you need to use Avogadro's number as a conversion factor.

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

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

PRACTICE Problems

Extra Practice Page 981 and glencoe.com

1. Zinc (Zn) is used to form a corrosion-inhibiting surface on galvanized steel. Determine the number of Zn atoms in 2.50 mol of Zn.
2. Calculate the number of molecules in 11.5 mol of water (H_2O).
3. Silver nitrate ($AgNO_3$) is used to make several different silver halides used in photographic films. How many formula units of $AgNO_3$ are there in 3.25 mol of $AgNO_3$?
4. **Challenge** Calculate the number of oxygen atoms in 5.0 mol of oxygen molecules. Oxygen is a diatomic molecule, O_2.

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

$$\begin{array}{c}\text{number of}\\ \text{representative particles}\end{array} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ representative particles}} = \text{number of moles}$$

For example, if instead of knowing how many moles of sucrose you have, suppose you knew that a sample contained 2.11×10^{24} molecules of sucrose. To convert this number of molecules of sucrose to moles of sucrose, you need a conversion factor that has moles in the numerator and molecules in the denominator.

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

Thus, 2.11×10^{24} molecules of sucrose is 3.50 mol of sucrose.

You can convert between moles and number of representative particles by multiplying the known quantity by the proper conversion factor. Example Problem 10.1 further illustrates the conversion process.

Reading Check **List** the two conversion factors that can be written from Avogadro's number.

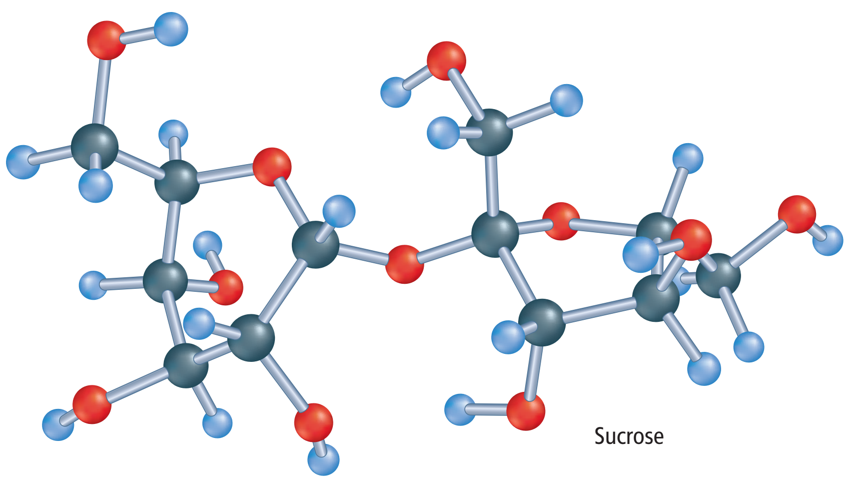


Figure 10.4 The representative particle of sucrose is a molecule. The ball-and-stick model shows that a molecule of sucrose is a single unit made up of carbon, hydrogen, and oxygen.
Analyze *Use the ball-and-stick model of sucrose to write the chemical formula for sucrose.*

EXAMPLE Problem 10.1

Math Handbook
Scientific Notation
pages 946–947

Converting Particles to Moles Zinc (Zn) 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 of zinc that contain 4.50×10^{24} atoms.

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 1 mol, you can predict that the answer should be less than 10 mol.

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

Unknown
moles Zn = ? mol

2 Solve for the Unknown

Use a conversion factor—the inverse of Avogadro's number—that relates moles to atoms.

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

Apply the conversion factor.

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

Substitute number of Zn atoms = 4.50×10^{24}. Multiply and divide numbers and units.

3 Evaluate the Answer

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

PRACTICE Problems

Extra Practice Page 981 and glencoe.com

5. How many moles contain each of the following?

a. 5.75×10^{24} atoms Al **b.** 2.50×10^{20} atoms Fe

6. Challenge Identify the representative particle for each formula, and convert the given number of representative particles to moles.

a. 3.75×10^{24} CO_2 **b.** 3.58×10^{23} $ZnCl_2$

Section 10.1 Assessment

Section Summary

- The mole is a unit used to count particles of matter indirectly. One mole of a pure substance contains Avogadro's number of representative particles.
- Representative particles include atoms, ions, molecules, formula units, electrons, and other similar particles.
- One mole of carbon-12 atoms has a mass of exactly 12 g.
- Conversion factors written from Avogadro's relationship can be used to convert between moles and number of representative particles.

7. MAIN Idea **Explain** why chemists use the mole.

8. State the mathematical relationship between Avogadro's number and 1 mol.

9. List the conversion factors used to convert between particles and moles.

10. Explain how a mole is similar to a dozen.

11. Apply How does a chemist count the number of particles in a given number of moles of a substance?

12. Calculate the mass of 0.25 mol of carbon-12 atoms.

13. Calculate the number of representative particles of each substance.

a. 11.5 mol Ag **c.** 0.150 mol NaCl
b. 18.0 mol H_2O **d.** 1.35×10^{-2} mol CCH_4

14. Arrange these three samples from smallest to largest in terms of number of representative particles: 1.25×10^{25} atoms of zinc (Zn), 3.56 mol of iron (Fe), and 6.78×10^{22} molecules of glucose ($C_6H_{12}O_6$).

Chemistry Online Self-Check Quiz glencoe.com

Section 10.2

Objectives

- **Relate** the mass of an atom to the mass of a mole of atoms.
- **Convert** between number of moles and the mass of an element.
- **Convert** between number of moles and number of atoms of an element.

Review Vocabulary

conversion factor: a ratio of equivalent values used to express the same quantity in different units

New Vocabulary

molar mass

Mass and the Mole

MAIN Idea A mole always contains the same number of particles; however, moles of different substances have different masses.

Real-World Reading Link When purchasing a dozen eggs, you can pick from several sizes—medium, large, and extra-large. The size of the egg does not affect how many come in the carton. A similar situation exists with the size of the atoms that make up a mole.

The Mass of a Mole

You would not expect a dozen limes to have the same mass as a dozen eggs. Because eggs and limes differ in size and composition, it is not surprising that they have different masses, as shown in **Figure 10.5.** One-mole quantities of two different substances have different masses for the same reason—the substances have different compositions. For example, if you put one mole of carbon and one mole of copper on separate balances, you would see a difference in mass, just as you do for the eggs and the limes. This occurs because carbon atoms differ from copper atoms. Thus, the mass of 6.02×10^{23} carbon atoms does not equal the mass of 6.02×10^{23} copper atoms.

Recall from Chapter 4 that each atom of carbon-12 has a mass of 12 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 approximately 1 amu, one-twelfth the mass of a carbon-12 atom. The mass of an atom of helium-4 is approximately 4 amu, one-third the mass of one atom of carbon-12.

You might have noticed, however, that the atomic-mass values given on the periodic table are not exact integers. For example, you will find 12.011 amu for carbon, 1.008 amu for hydrogen, and 4.003 amu for helium. These noninteger values occur because the values are weighted averages of the masses of all the naturally occurring isotopes of each element.

Figure 10.5 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.

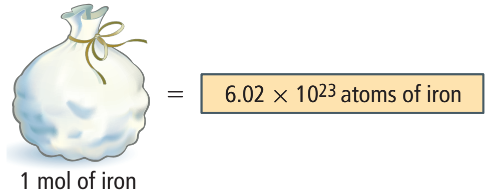
Matt Meadows

■ **Figure 10.6** One mole of iron, represented by a bag of particles, contains Avogadro's number of atoms and has a mass equal to its atomic mass in grams.
Apply ***What is the mass of one mole of copper?***

Concepts In Motion

Interactive Figure To see an animation of molar mass, visit glencoe.com.

CLICK HERE



Molar Mass How does the mass of one atom relate to the mass of one mole of that atom? Recall that the mole is defined as the number of carbon-12 atoms in exactly 12 g of pure carbon-12. Thus, the mass of one mole of carbon-12 atoms is 12 g. 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 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. As given on the periodic table, an atom of iron has an atomic mass of 55.845 amu. Thus, the molar mass of iron is 54.845 g/mol, and 1 mol (or 6.02×10^{23} atoms of iron) has a mass of 55.845 g. Note that by measuring 55.845 g of iron, you indirectly count out 6.02×10^{23} atoms of iron. **Figure 10.6** shows the relationship between molar mass and one mole of an element.

PROBLEM-SOLVING LAB

Formulate a Model

How are molar mass, Avogadro's number, and the atomic nucleus related?
A nuclear model of mass can provide a simple picture of the connections among the mole, molar mass, and the number of representative particles in a mole.

Analysis

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

Think Critically

1. **Apply** 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. **Draw** Carbon-12 contains six protons and six neutrons. Draw the carbon-12 nucleus and calculate the mass of one atom in amu and g.
3. **Apply** 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.
4. **Apply** 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?
5. **Conclude** What can you conclude about the relationship between the number of atoms and the mass of each sample?

Matt Meadows

Using Molar Mass

Imagine that your class bought jelly beans in bulk to sell by the dozen at a candy sale. You soon realize that it is too much work to count out each dozen, so you instead decide to measure the jelly beans by mass. You find that the mass of 1 dozen jelly beans is 35 g. This relationship and the conversion factors that stem from it are as follows:

$$1 \text{ dozen jelly beans} = 35 \text{ g jelly beans}$$

$$\frac{35 \text{ g jelly beans}}{1 \text{ dozen jelly beans}} \text{ and } \frac{1 \text{ dozen jelly beans}}{35 \text{ g jelly beans}}$$

What mass of jelly beans should you measure if a customer wants 5 dozen jelly beans? To determine this mass, you would multiply the number of dozens of jelly beans to be sold by the correct conversion factor. Select the conversion factor with the units you are converting to in the numerator (g) and the units you are converting from in the denominator (dozen).

$$5 \cancel{\text{ dozen jelly beans}} \times \frac{35 \text{ g jelly beans}}{1 \cancel{\text{ dozen jelly beans}}} = 175 \text{ g jelly beans.}$$

A quantity of 5 dozen jelly beans has a mass of 175 g.

Reading Check **Compare** How are the jelly bean conversion factors used above similar to the molar mass of a compound?

Chemistry Online
Personal Tutor For an online tutorial on using conversion factors, visit glencoe.com.

Moles to mass Now suppose that while working in a chemistry lab, you need 3.00 mol of copper (Cu) for a chemical reaction. How would you measure that amount? Like the 5 dozen jelly beans, the number of moles of copper can be converted to an equivalent mass and measured on a balance.

To calculate the mass of a given number of moles, simply multiply the number of moles by the molar mass.

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

If you check the periodic table, you will find that copper, element 29, has an atomic mass of 63.546 amu. You know that the molar mass of an element (in g/mol) is equal to its atomic mass (given in amu). Thus, copper has a molar mass of 63.546 g/mol. By using the molar mass, you can convert 3.00 mol of copper to grams of copper.

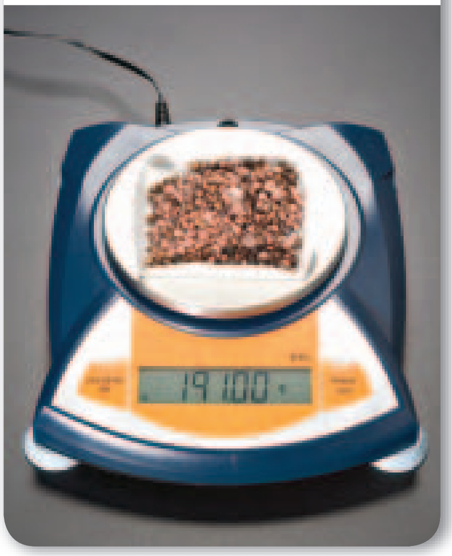
$$3.00 \cancel{\text{ mol Cu}} \times \frac{63.546 \text{ g Cu}}{1 \cancel{\text{ mol Cu}}} = 191 \text{ g Cu}$$

So, as shown in **Figure 10.7,** you can measure the 3.00 mol of copper needed for the reaction by using a balance to measure out 191 g of copper. 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?

Connection to Biology Cellular biologists continually discover new biologic proteins. After a new biomolecule is discovered, biologists determine the molar mass of the compound using a technique known as mass spectrometry. In addition to the molar mass, mass spectrometry also provides additional information that helps the biologist reveal the compound's composition.

FOLDABLES
Incorporate information from this section into your Foldable.

Figure 10.7 To measure 3.00 mol of copper, place a weighing paper on a balance, tare the balance, and then add the 191 g of copper filings.



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EXAMPLE Problem 10.2

Math Handbook
Rounding
page 952

Mole-to-Mass Conversion Chromium (Cr), a transition element, is a component of chrome plating. Chrome plating is used on metals and in steel alloys to control corrosion. Calculate the mass in grams of 0.0450 mol Cr.

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 of a mole, the answer should be less than one-tenth of the molar mass.

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

Unknown
mass Cr = ? g

2 Solve for the Unknown

Use a conversion factor—the molar mass—that relates grams of chromium to moles of chromium. Write the conversion factor with moles of chromium in the denominator and grams of chromium in the numerator. Substitute the known values into the equation and solve.

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

Apply the conversion factor.

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

Substitute 0.450 mol for moles Cr and 52.00 g/mol for molar mass of Cr. Multiply and divide number and units.

3 Evaluate the Answer

The known number of moles of chromium has the smallest number of significant figures, three, so the answer is correctly stated with three digits. The answer is less than one-tenth the mass of 1 mol, as predicted, and is in grams, a mass unit.

PRACTICE Problems

Extra Practice Page 981 and glencoe.com

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

a. 3.57 mol Al

b. 42.6 mol Si

16. Challenge Convert each given quantity in scientific notation to mass in grams expressed in scientific notation.

a. 3.45×10^{2} mol Co

b. 2.45×10^{-2} mol Zn

Real-World Chemistry
The Importance of Chromium

Chromium What gives these rims their mirrorlike finish? The metal alloy rim has been plated, or coated, with a thin layer of chromium. Chrome plating has been used in the automobile industry for decades because of its beauty and its corrosion resistance.

If you examine the atomic mass values given on the periodic table, you will notice that the values differ in their number of significant figures; most atomic mass values have four or five significant figures. When you use an atomic mass value from the periodic table, use all the significant figures provided. If your calculation involves several steps, do not round answers until the end of the calculation. By doing this, you increase the precision of any calculation involving atomic mass.

EXAMPLE Problem 10.3

Math Handbook
Dimensional Analysis
page 956

Mass-to-Mole Conversion Calcium (Ca), 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 Ca?

1 Analyze the Problem

You must convert the mass of calcium 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 10 mol.

Known
mass = 525 g Ca
molar mass Ca = 40.08 g/mol Ca

Unknown
number of moles Ca = ? mol

2 Solve for the Unknown

Use a conversion factor—the inverse of molar mass—that relates moles of calcium to grams of calcium. Substitute the known values and solve.

$$\text{mass Ca} \times \frac{1 \text{ mol Ca}}{\text{grams Ca}} = \textbf{moles Ca}$$

Apply the conversion factor.

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

Substitute mass Ca = 525 g, and inverse molar mass of Ca = 1 mol/40.08 g. Multiply and divide numbers and units.

3 Evaluate the Answer

The mass of calcium has the fewest significant figures, three, so the answer is expressed correctly with three digits. As predicted, the answer is greater than 10 mol and has the expected unit.

PRACTICE Problems

Extra Practice Page 981 and glencoe.com

17. Determine the number of moles in each of the following.
a. 25.5 g Ag **b.** 300.0 g S

18. Challenge Convert each mass to moles. Express the answer in scientific notation.
a. 1.25×10^3 g Zn **b.** 1.00 kg Fe

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

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

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

The conversion factor relating number of jelly beans and dozens is, 12 jelly beans/dozen. Applying it yields the answer in jelly beans.

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

The 550 g of leftover jelly beans is equal to 192 jelly beans.

Just as you could not make a direct conversion from the mass of jelly beans to the number of jelly beans, you cannot make a direct conversion from the mass of a substance to the number of representative particles of that substance. You must first convert mass to moles by multiplying by a conversion factor that relates moles and mass. That conversion factor is the molar mass. The number of moles must then be multiplied by a conversion factor that relates the number of representative particles to moles. For this conversion, you use will use Avogadro's number. This two-step process is shown in Example Problem 10.4.

EXAMPLE Problem 10.4

Math Handbook
Dimensional Analysis
page 956

Mass-to-Atoms Conversion Gold (Au) is one of a group of metals called the coinage metals (copper, silver, and gold). How many atoms of gold are in a U.S. Eagle, a gold alloy bullion coin with a mass of 31.1 g Au?

1 Analyze the Problem

You must determine the number of atoms in a given mass of gold. Because you cannot convert directly from mass to the number of atoms, you must first convert the mass to moles using the molar mass. Then, convert moles to the number of atoms using Avogadro's number. The given mass of the gold coin is about one-sixth the molar mass of gold (196.97 g/mol), so the number of gold atoms should be approximately one-sixth Avogadro's number.

Known

mass = 31.1 g Au

molar mass Au = 196.97 g/mol Au

Unknown

number of atoms Au = ?

2 Solve for the Unknown

Use a conversion factor—the inverse of the molar mass—that relates moles of gold to grams of gold.

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

Apply the conversion factor.

$$31.1 \cancel{\text{g Au}} \times \frac{1 \text{ mol Au}}{196.97 \cancel{\text{g Au}}} = 0.158 \text{ mol Au}$$

Substitute mass Au = 31.1 g and the inverse molar mass of Au = 1 mol/196.97 g. Multiply and divide numbers and units.

To convert the calculated moles of gold to atoms, multiply by Avogadro's number.

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

Apply the conversion factor.

$$0.158 \cancel{\text{mol Au}} \times \frac{6.02 \times 10^{23} \text{ atoms Au}}{1 \cancel{\text{mol Au}}} = \mathbf{9.51 \times 10^{22} \text{ atoms Au}}$$

Substitute moles Au = 0.158 mol, and solve.

3 Evaluate the Answer

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

EXAMPLE Problem 10.5

Math Handbook
Calculations with Significant Figures
pages 952–953

Atoms-to-Mass Conversion Helium (He) 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 liquefied. A party balloon contains 5.50×10^{22} atoms of helium 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. First, convert the number of atoms to moles, then convert moles to grams.

Known

number of atoms He = 5.50×10^{22} atoms He

molar mass He = 4.00 g/mol He

Unknown

mass = ? g He

2 Solve for the Unknown

Use a conversion factor—the inverse of Avogadro's number—that relates moles to number of atoms.

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

Apply the conversion factor.

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

Substitute atoms He = 5.50×10^{22} atoms. Multiply and divide numbers and units.

Next, apply a conversion factor—the molar mass of helium—that relates mass of helium to moles of helium.

$$\text{moles He} \times \frac{\text{grams He}}{1 \text{ mol He}} = \textbf{mass He}$$

Apply the conversion factor.

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

Substitute moles He = 0.0914 mol, molar mass He = 4.00 g/mol, and solve.

3 Evaluate the Answer

The answer is expressed correctly with three significant figures and is in grams, a mass unit.

PRACTICE Problems

Extra Practice Page 981 and glencoe.com

19. 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

20. 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

21. Challenge Convert each given mass to number of representative particles. Identify the type of representative particle, and express the number in scientific notation.

a. 4.56×10^{3} g Si

b. 0.120 kg Ti

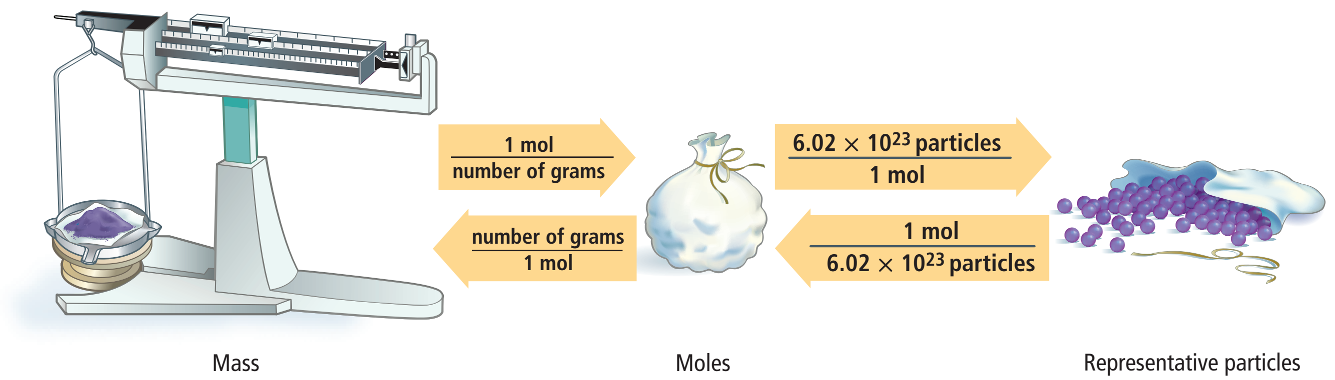


Figure 10.8 The mole is at the center of conversions between mass and particles (atoms, ions, or molecules). In the figure, mass is represented by a balance, moles by a bag of particles, and representative particles by the contents that are spilling out of the bag. Two steps are needed to convert from mass to representative particles or the reverse.

Now that you have practiced conversions between mass, moles, and representative particles, you probably realize 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 10.8** shows the steps to follow as you complete these conversions. In the Example Problems, two steps were used to convert either mass to moles to atoms, or atoms to moles to mass. Instead of two separate steps, these conversions can be made in one step. Suppose you want to find out how many atoms of oxygen are in 1.00 g of oxygen. This calculation involves two conversions—mass to moles and then moles to atoms. You could set up one equation like this.

$$1.00\ \cancel{\text{g O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{31.998\ \cancel{\text{g O}_2}} \times \frac{6.02 \times 10^{23}\ \text{atoms O}_2}{1\ \cancel{\text{mol O}_2}}$$
$$= 1.88 \times 10^{22}\ \text{atoms O}_2$$

Section 10.2 Assessment

Section Summary

- The mass in grams of one mole of any pure substance is called its molar mass.
- The molar mass of an element is numerically equal to its atomic mass.
- 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 to mass. The inverse of molar mass is used to convert from mass to moles.

22. MAIN Idea **Summarize** in terms of particles and mass, one-mole quantities of two different monatomic elements.
23. **State** the conversion factors needed to convert between mass and moles of the element fluorine.
24. **Explain** how molar mass relates the mass of an atom to the mass of a mole of atoms.
25. **Describe** the steps used to convert the mass of an element to the number of atoms of the element.
26. **Arrange** these quantities from smallest to largest in terms of mass: 1.0 mol of Ar, 3.0×10^{24} atoms of Ne, and 20 g of Kr.
27. **Identify** the quantity that is calculated by dividing the molar mass of an element by Avogadro's number.
28. **Design** a concept map that shows the conversion factors needed to convert between mass, moles, and number of particles.

Chemistry Online Self-Check Quiz glencoe.com

Section 10.3

Objectives

- **Recognize** the mole relationships shown by a chemical formula.
- **Calculate** the molar mass of a compound.
- **Convert** between the number of moles and mass of a compound.
- **Apply** conversion factors to determine the number of atoms or ions in a known mass of a compound.

Review Vocabulary

representative particle: an atom, molecule, formula unit, or ion

Moles of Compounds

MAIN Idea The molar mass of a compound can be calculated from its chemical formula and can be used to convert from mass to moles of that compound.

Real-World Reading Link Imagine checking two pieces of luggage at the airport, only to find out that one of them is over the weight limit. Because the weight of each suitcase depends on the combination of the items packed inside, changing the combination of the items in the two suitcases changes the weight of each.

Chemical Formulas and the Mole

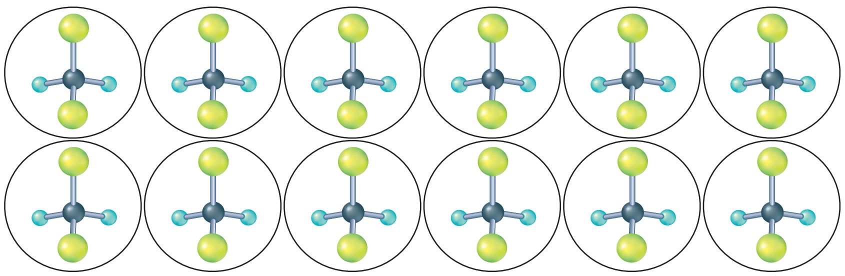
You have learned that different kinds of representative particles are counted using the mole. In the last section, you read how to use molar mass to convert among moles, mass, and number of particles of an element. Can you make similar conversions for compounds and ions? Yes, you can, but to do so you will need to know the molar mass of the compounds and ions involved.

Recall that a chemical formula indicates the numbers and types of atoms contained in one unit of the compound. Consider the compound dichlorodifluoromethane with the chemical formula CCl_2F_2. The subscripts in the formula indicate that one molecule of CCl_2F_2 consists of one carbon (C) atom, two chlorine (Cl) atoms, and two fluorine (F) atoms. These atom are chemically bonded together. The C-Cl-F ratio in CCl_2F_2 is 1:2:2.

Now suppose you had a mole of CCl_2F_2. The representative particles of the compound are molecules, and a mole of CCl_2F_2 contains Avogadro's number of molecules. The C-Cl-F ratio in one mole of CCl_2F_2 would still be 1:2:2, as it is in one molecule the compound. **Figure 10.9** illustrates this principle for a dozen CCl_2F_2 molecules. Check for yourself that a dozen CCl_2F_2 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 CCl_2F_2, it also represents a mole of the compound.

Figure 10.9 A dozen freon molecules contains one dozen carbon atoms, two dozen chlorine atoms, and two dozen fluorine atoms.

Interpret *How many of each kind of atom—carbon, chlorine, and fluorine—are contained in 1 mol of CCl_2F_2?*



VOCABULARY

ACADEMIC VOCABULARY

Ratio

the relationship in size or quantity of two or more things; proportion

The test results showed his LDL-to-HDL cholesterol ratio was too high.

In some chemical calculations, you might need to convert between moles of a compound and moles of individual atoms in the compound. The following ratios, or conversion factors, can be written for use in these calculations for the molecule CCl_2F_2.

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

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

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

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

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 10.6

Math Handbook

Dimensional Analysis page 956

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

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 2 mol of Al^{3+} ions. Thus, the answer should be two times the number of moles of Al_2O_3.

Known	**Unknown**
number of moles = 1.25 mol Al_2O_3	**number of moles = ? mol Al^{3+} ions**

2 Solve for the Unknown

Use the relationship that 1 mol of Al_2O_3 contains 2 mol of Al^{3+} ions to write a conversion factor.

$$\frac{2 \text{ mol } Al^{3+} \text{ ions}}{1 \text{ mol } Al_2O_3}$$ **Create a conversion factor relating moles of Al^{3+} ions to moles of Al_2O_3.**

To convert the known number of moles of Al_2O_3 to moles of Al^{3+} ions, multiply by the ions-to-moles conversion factor.

$$\text{moles } Al_2O_3 \times \frac{2 \text{ mol } Al^{3+} \text{ ions}}{1 \text{ mol } Al_2O_3} = \textbf{moles } \mathbf{Al^{3+}} \textbf{ ions}$$ **Apply the conversion factor.**

$$1.25 \cancel{\text{mol } Al_2O_3} \times \frac{2 \text{ mol } Al^{3+} \text{ ions}}{1 \cancel{\text{mol } Al_2O_3}} = \mathbf{2.50 \text{ mol } Al^{3+} \text{ ions}}$$ **Substitute moles Al_2O_3 = 1.25 mol Al_2O_3 and solve.**

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.

PRACTICE Problems

Extra Practice Pages 981–982 and glencoe.com

29. Zinc chloride ($ZnCl_2$) is used in soldering flux, an alloy used to join two metals together. Determine the moles of Cl^- ions in 2.50 mol $ZnCl_2$.

30. Plants and animals depend on glucose ($C_6H_{12}O_6$) as an energy source. Calculate the number of moles of each element in 1.25 mol $C_6H_{12}O_6$.

31. Iron(III) sulfate[$Fe_2(SO_4)_3$] is sometimes used in the water purification process. Determine the number of moles of sulfate ions present in 3.00 mol of $Fe_2(SO_4)_3$.

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

33. Challenge Calculate the number of moles of hydrogen atoms in 1.15×10^1 mol of water. Express the answer in scientific notation.

The Molar Mass of Compounds

The mass of your backpack is the sum of the mass of the pack and 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 all the particles that make up the compound.

Suppose you want to determine the molar mass of the compound potassium chromate (K_2CrO_4). Start by looking up the molar mass of each element present in K_2CrO_4. Then, multiply each molar mass by the number of moles of that element in the chemical formula. Adding the masses of each element yields the molar mass of K_2CrO_4.

$$2\ \cancel{\text{mol K}} \times \frac{39.10\ \text{g K}}{1\ \cancel{\text{mol K}}} = 78.20\ \text{g}$$

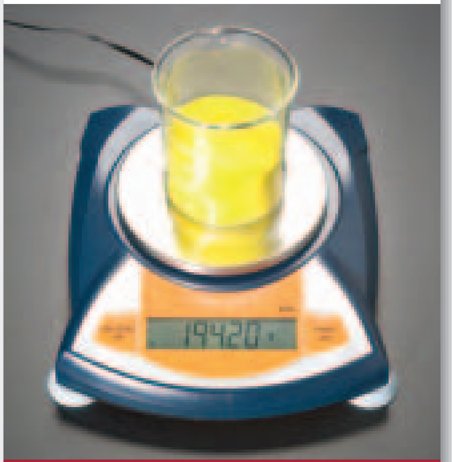
$$1\ \cancel{\text{mol Cr}} \times \frac{52.00\ \text{g Cr}}{1\ \cancel{\text{mol Cr}}} = 52.00\ \text{g}$$

$$4\ \cancel{\text{mol O}} \times \frac{16.00\ \text{g O}}{1\ \cancel{\text{mol O}}} = 64.00\ \text{g}$$

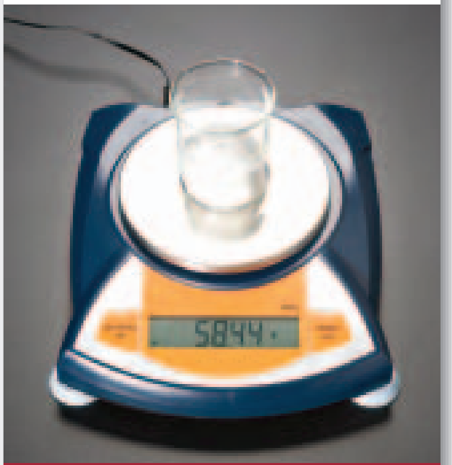
$$\text{molar mass } K_2CrO_4 = 194.20\ \text{g}$$

The molar mass of a compound demonstrates the law of conservation of mass; the total mass of the reactants that reacted equals the mass of the compound formed. **Figure 10.10** shows equivalent masses of one mole of potassium chromate, sodium chloride, and sucrose.

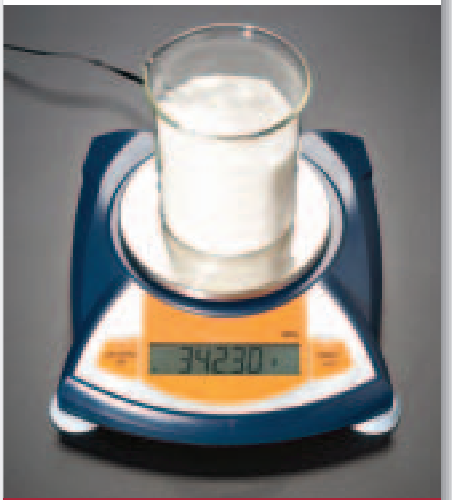
■ **Figure 10.10** Because each substance contains different numbers and kinds of atoms, their molar masses are different. The molar mass of each compound is the sum of the masses of all the elements contained in the compound.



Potassium chromate (K_2CrO_4)



Sodium chloride (NaCl)



Sucrose ($C_{12}H_{22}O_{11}$)

PRACTICE Problems

Extra Practice Pages 981–982 and glencoe.com

34. Determine the molar mass of each ionic compound.

a. NaOH **b.** $CaCl_2$ **c.** $KC_2H_3O_2$

35. Calculate the molar mass of each molecular compound.

a. C_2H_5OH **b.** HCN **c.** CCl_4

36. Challenge Identify each substance as a molecular compound or an ionic compound, and then calculate its molar mass.

a. $Sr(NO_3)_2$ **b.** $(NH_4)_3PO_4$ **c.** $C_{12}H_{22}O_{11}$

Matt Meadows

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, you can measure that mass on a balance. In Example Problem 10.2, 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 10.7

Math Handbook
Calculations with Significant Figures
pages 952–953

Mole-to-Mass Conversion for Compounds The characteristic odor of garlic is due to allyl sulfide [$(C_3H_5)_2S$]. What is the mass of 2.50 mol of $(C_3H_5)_2S$?

1 Analyze the Problem

You are given 2.50 mol of $(C_3H_5)_2S$ 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 = ? g/mol $(C_3H_5)_2S$
mass = ? g $(C_3H_5)_2S$

2 Solve for the Unknown

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

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

Multiply the moles of S in the compound by the molar mass of S.

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

Multiply the moles of C in the compound by the molar mass of C.

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

Multiply the moles of H in the compound by the molar mass of H.

molar mass = (32.07 g + 72.06 g + 10.08 g) = **114.21 g/mol $(C_3H_5)_2S$**

Total the mass values.

Use a conversion factor—the molar mass—that relates grams to moles.

$$\text{moles } (C_3H_5)_2S \times \frac{\text{grams } (C_3H_5)_2S}{1\ \text{mol } (C_3H_5)_2S} = \textbf{mass } (C_3H_5)_2S$$

Apply the conversion factor.

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

Substitute moles $(C_3H_5)_2S$ = 2.5 mol, molar mass $(C_3H_5)_2S$ = 114.21 g/mol, and solve.

PRACTICE Problems

Extra Practice Pages 981–982 and glencoe.com

37. The United States chemical industry produces more sulfuric acid (H_2SO_4), in terms of mass, than any other chemical. What is the mass of 3.25 mol of H_2SO_4?

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

39. **Challenge** Write the chemical formula for potassium permanganate, and then calculate the mass in grams of 2.55 mol of the compound.

Converting the Mass of a Compound to Moles

Imagine that an 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 10.8

Math Handbook
Calculations with Significant Figures
pages 952–953

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 of the compound.

1 Analyze the Problem

You are given 325 g of $Ca(OH)_2$ and must solve 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 $Ca(OH)_2$.

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

Multiply the moles of Ca in the compound by the molar mass of Ca.

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

Multiply the moles of O in the compound by the molar mass of O.

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

Multiply the moles of H in the compound by the molar mass of H.

molar mass = (40.08 g + 32.00 g + 2.016 g) = **74.10 g/mol $Ca(OH)_2$** **Total the mass values.**

Use a conversion factor—the inverse of the molar mass—that relates moles to grams.

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

Apply the conversion factor. Substitute mass Ca = 325 g, inverse molar mass $Ca(OH)_2$ = 1 mol/74.10 g, and solve.

3 Evaluate the Answer

To check the reasonableness of the answer, round 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. The unit, moles, is correct, and there are three significant figures.

PRACTICE Problems

Extra Practice Pages 981–982 and glencoe.com

40. Determine the number of moles present in each compound.

a. 22.6 g $AgNO_3$ **b.** 6.50 g $ZnSO_4$ **c.** 35.0 g HCl

41. Challenge Identify each as an ionic or molecular compound and convert the given mass to moles. Express your answers in scientific notation.

a. 2.50 kg Fe_2O_3 **b.** 25.4 mg $PbCl_4$

FOLDABLES Incorporate information from this section into your Foldable.

Converting the Mass of a Compound to Number of Particles

Example Problem 10.8 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 1 mol of compound. These are based on the chemical formula. Example Problem 10.9 provides practice in solving this type of problem.

EXAMPLE Problem 10.9

Math Handbook
Calculations with Significant Figures
pages 952–953

Conversion from Mass to Moles to Particles Aluminum chloride ($AlCl_3$) is used in refining petroleum and manufacturing rubber and lubricants. A sample of aluminum chloride 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?

1 Analyze the Problem

You are given 35.6 g of $AlCl_3$ 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 same ratio. The mass of one formula unit in grams will 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 $AlCl_3$.

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

Multiply the moles of Al in the compound by the molar mass of Al.

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

Multiply the moles of Cl in the compound by the molar mass of Cl.

$$\text{molar mass} = (26.98\ \text{g} + 106.35\ \text{g}) = 133.33\ \text{g/mol AlCl}_3$$

Total the molar mass values.

Use a conversion factor—the inverse of the molar mass—that relates moles to grams.

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

Apply the conversion factor.

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

Substitute mass $AlCl_3$ = 35.6 g and inverse molar mass $AlCl_3$ = 1 mol/133.33 g, and solve.

Use Avogadro's number.

$0.267 \cancel{\text{mol AlCl}_3} \times \frac{6.02 \times 10^{23} \text{ formula units}}{1 \cancel{\text{mol AlCl}_3}}$ **Multiply and divide numbers and units.**

$= 1.61 \times 10^{23}$ formula units $AlCl_3$

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

$1.61 \times 10^{23} \cancel{\text{AlCl}_3 \text{ formula units}} \times \frac{1 \text{ Al}^{3+} \text{ ion}}{1 \cancel{\text{AlCl}_3 \text{ formula unit}}}$ **Multiply and divide numbers and units.**

$= \mathbf{1.61 \times 10^{23}\ Al^{3+}\ ions}$

$1.61 \times 10^{23} \cancel{\text{AlCl}_3 \text{ formula units}} \times \frac{3 \text{ Cl}^{-} \text{ ions}}{1 \cancel{\text{AlCl}_3 \text{ formula unit}}}$ **Multiply and divide numbers and units.**

$= \mathbf{4.83 \times 10^{23}\ Cl^{-}\ ions}$

Calculate the mass in grams of one formula unit of $AlCl_3$. Use the inverse of Avogadro's number as a conversion factor.

$\frac{133.33 \text{ g AlCl}_3}{1 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}}}{6.02 \times 10^{23} \text{ formula units}}$ **Substitute mass $AlCl_3$ = 133.33 g, and solve.**

$= \mathbf{2.21 \times 10^{-22}\ g\ AlCl_3/formula\ unit}$

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

Extra Practice Pages 981–982 and glencoe.com

42. Ethanol (C_2H_5OH), a domestically produced fuel source, is often blended with gasoline. A sample of ethanol 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?

43. 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?

44. A sample of carbon dioxide (CO_2) 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?

45. What mass of sodium chloride (NaCl) contains 4.59×10^{24} formula units?

46. Challenge A sample of silver chromate has a mass of 25.8 g.

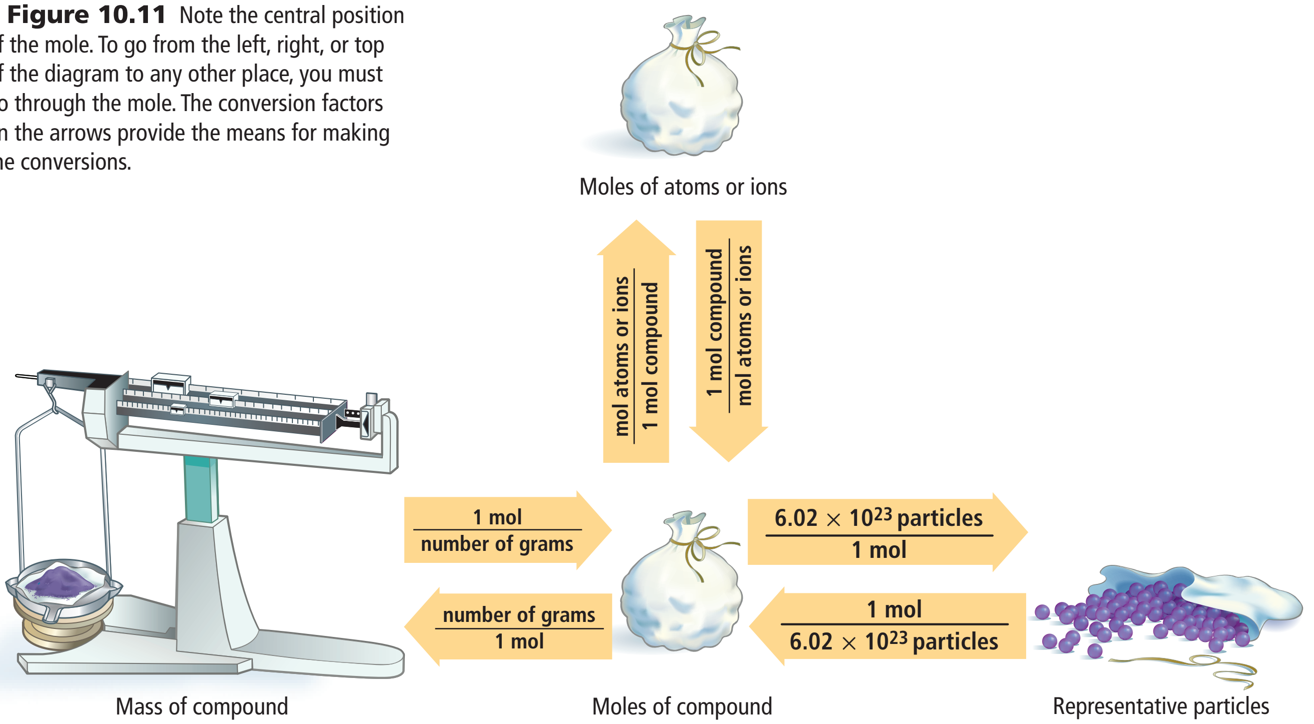
a. Write the formula for silver chromate.

b. How many cations are present in the sample?

c. How many anions are present in the sample?

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

Figure 10.11 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 between mass, moles, and the number of particles are summarized in **Figure 10.11.** Note that molar mass and the inverse of molar mass are conversion factors between mass and number of moles. Avogadro's number and its inverse are the conversion factors between moles and the number of representative particles. To convert between moles and the number of moles of atoms or ions contained in the compound, use 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 10.11.** These ratios are derived from the subscripts in the chemical formula.

Section 10.3 Assessment

Section Summary

- Subscripts in a chemical formula indicate how many moles of each element are present in 1 mol of the compound.
- The molar mass of a compound is calculated from the molar masses of all the elements in the compound.
- Conversion factors based on a compound's molar mass are used to convert between moles and mass of a compound.

47. MAIN Idea **Describe** how to determine the molar mass of a compound.
48. **Identify** the conversion factors needed to convert between the number of moles and the mass of a compound.
49. **Explain** how you can determine the number of atoms or ions in a given mass of a compound.
50. **Apply** How many moles of K, C, and O atoms are there in 1 mol of $K_2C_2O_4$?
51. **Calculate** the molar mass of $MgBr_2$.
52. **Calculate** Calcium carbonate is the calcium source for many vitamin tablets. The recommended daily allowance of calcium is 1000 mg of Ca^{2+} ions. How many moles of Ca^{2+} does 1000 mg represent?
53. **Design** a bar graph that will show the number of moles of each element present in 500 g of a particular form of dioxin ($C_{12}H_4Cl_4O_2$), a powerful poison.

Chemistry Online Self-Check Quiz glencoe.com

Section 10.4

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.

Review Vocabulary

percent by mass: the ratio of the mass of each element to the total mass of the compound expressed as a percent

New Vocabulary

percent composition
empirical formula
molecular formula

Empirical and Molecular Formulas

MAIN Idea A molecular formula of a compound is a whole-number multiple of its empirical formula.

Real-World Reading Link You might have noticed that some beverage bottles and food packages contain two or more servings instead of the single serving you expect. How would you determine the total number of calories contained in the package?

Percent Composition

Chemists, such as those shown in **Figure 10.12,** 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.

It is the analytical chemist's job to identify the elements a compound contains and determine their percents by mass. Gravimetric and volumetric analyses are experimental procedures based on the measurement of mass for solids and liquids, respectively.

Percent composition from experimental data For example, consider a 100-g sample of a compound that 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.

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

Figure 10.12 New compounds are first made on a small scale by a synthetic chemist like the one shown on the left. Then, an analytical chemist, like the one shown on the right, analyzes the compound to verify its structure and percent composition.

CAREERS IN CHEMISTRY

Medicinal Chemist Much like a chef trying to perfect a recipe, the medicinal chemist works to perfect the best combination of reactants. They use their knowledge of the effects of toxins and medicines on the human body to synthesize new molecules that target disease. For more information on chemistry careers, visit **glencoe.com**.

Because percent means parts per 100, the percents by mass of all the elements of a compound must always add up to 100.

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

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

Thus, the compound is 55% X and 45% Y. The percent by mass of each element in a compound is the **percent composition** of a compound.

Percent composition from the chemical formula The percent composition of a compound can also be determined from its chemical formula. To do this, assume you have exactly 1 mol of the compound and use the chemical formula to calculate the compound's molar mass. Then, determine the mass of each element in a mole of the compound by multiplying the element's molar mass by its subscript in the chemical formula. Finally, use the equation below to find the percent by mass of each element.

Percent by Mass from the Chemical Formula

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

The percent by mass of an element in a compound is the mass of the element in 1 mol of the compound divided by the molar mass of the compound, multiplied by 100.

Example Problem 10.10 covers calculating percent composition.

MiniLab

Analyze Chewing Gum

Are sweetening and flavoring added as a coating or mixed throughout chewing gum?

Procedure

1. Read and complete the lab safety form.
2. Unwrap two pieces of **chewing gum.** Place each piece on a **weighing paper.** Measure and record each mass using a **balance.**
   WARNING: *Do not eat any items used in the lab.*
3. Add 150 mL of cold **tap water** to a **250-mL beaker.** Place one piece of chewing gum in the water, and stir with a **stirring rod** for 2 min.
4. Pat the gum dry using **paper towels.** Measure and record the mass of the dried gum.
5. Use **scissors** to cut the second piece of gum into small pieces. Repeat Step 3 using fresh water. Keep the pieces from clumping together.
   WARNING: *Use caution with scissors.*
6. Use a 10-cm × 10-cm piece of **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.

Analysis

1. **Calculate** 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. **Calculate** For the gum cut into small pieces, calculate the mass of dissolved sweeteners and flavorings.
3. **Apply** For each piece of gum, determine the percent of the original mass from the soluble sweeteners and flavorings.
4. **Infer** What can you infer from the two percentages? Is the gum sugar-coated or are the sweeteners and flavorings mixed throughout?

EXAMPLE Problem 10.10

Math Handbook
Percents
page 965

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

1 Analyze the Problem

You are given only the chemical formula. Assume you have 1 mol of $NaHCO_3$. Calculate the molar mass and the mass of each element in 1 mol to determine the percent by mass of each element in the compound. The sum of all percents should be 100, although your answer might vary slightly due to rounding.

Known
formula = $NaHCO_3$

Unknown
percent Na = ?
percent H = ?
percent C = ?
percent O = ?

2 Solve for the Unknown

Determine the molar mass of $NaHCO_3$ and each element's contribution.

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

Multiply the molar mass of Na by the number of Na atoms in the compound.

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

Multiply the molar mass of H by the number of H atoms in the compound.

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

Multiply the molar mass of C by the number of C atoms in the compound.

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

Multiply the molar mass of O by the number of O atoms in the compound.

$$\text{molar mass} = (22.99 \text{ g} + 1.008 \text{ g} + 12.01 \text{ g} + 48.00 \text{ g})$$
$$= 84.01 \text{ g/mol NaHCO}_3$$

Total the mass values.

Use the percent by mass equation.

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

State the equation.

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

Substitute mass of Na in 1 mol compound = 22.99 g/mol and molar mass $NaHCO_3$ = 84.01 g/mol. Calculate % Na.

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

Substitute mass of H in 1 mol compound = 1.0089 g/mol and molar mass $NaHCO_3$ = 84.01 g/mol. Calculate % H.

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

Substitute mass of C in 1 mol compound = 12.01 g/mol and molar mass $NaHCO_3$ = 84.01 g/mol. Calculate % C.

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

Substitute mass of O in 1 mol compound = 48.00 g/mol and molar mass $NaHCO_3$ = 84.01 g/mol. Calculate % O.

$NaHCO_3$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 with four significant figures. When rounding error is accounted for, the sum of the mass percents is 100%, as required.

PRACTICE Problems

Extra Practice Page 982 and **glencoe.com**

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

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

56. Calcium chloride ($CaCl_2$) is sometimes used as a de-icer. Calculate the percent by mass of each element in $CaCl_2$.

57. Challenge Sodium sulfate is used in the manufacture of detergents.

a. Identify each of the component elements of sodium sulfate, and write the compound's chemical formula.

b. Identify the compound as ionic or covalent.

c. Calculate the percent by mass of each element in sodium sulfate.

Empirical Formula

When a compound's percent composition is known, its formula can be calculated. First, determine the smallest whole-number ratio of the moles of the elements in the compound. This ratio gives 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 might or might 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.

Percent composition or masses of the elements in a given mass of compound can be used to determine the formula for the compound. If percent composition is given, assume 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. This can be seen in **Figure 10.13,** where 100.00 g of the 40.05% S and 59.95% O compound contains 40.05 g of S and 59.95 g of O. The mass of each element is then converted to moles.

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

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

Thus, the mole ratio of S atoms to O atoms in the oxide is 1.249:3.747.

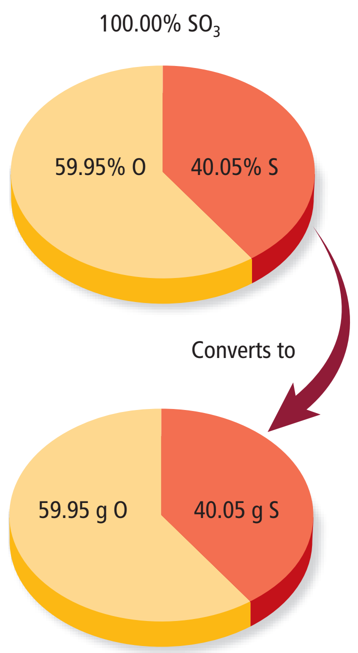
When the values in a mole ratio are not whole numbers, they cannot be used as subscripts in a chemical formula. You can convert the ratio to whole numbers by recognizing that the element with the smallest number of moles might have the smallest subscript possible, 1. To make the mole value of sulfur equal to 1, divide both mole values by the moles of sulfur (1.249). This does not change the ratio between the two elements because both are divided by the same number.

$$\frac{1.249\ \text{mol S}}{1.249} = 1\ \text{mol S} \qquad \frac{3.747\ \text{mol O}}{1.249} = 3\ \text{mol O}$$

The simplest whole-number mole ratio of S to O is 1:3. Thus, the empirical formula is SO_3. Sometimes, dividing by the smallest mole value does not yield whole numbers. In such cases, each mole value must then be multiplied by the smallest factor that will make it a whole number. This is shown in Example Problem 10.11.

Reading Check **List** the steps needed to calculate the empirical formula from percent composition data.

Figure 10.13 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.



EXAMPLE Problem 10.11

Math Handbook
Ratios
page 964

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, convert from grams to moles and find the smallest whole-number ratio of moles of the elements.

Known
percent by mass C = 48.64% C
percent by mass H = 8.16% H
percent by mass O = 43.20% O

Unknown
empirical formula = ?

2 Solve for the Unknown

Convert each mass to moles using a conversion factor—the inverse of the molar mass—that relates moles to grams.

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

Substitute mass C = 48.64 g, inverse molar mass C = 1 mol/12.01 g, and calculate moles of C.

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

Substitute mass H = 8.16 g, inverse molar mass H = 1 mol/1.008 g, and calculate moles of H.

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

Substitute mass O = 43.20 g, inverse molar mass O = 1 mol/16.00 g, and calculate moles of O.

Methyl acetate has a mole ratio of (4.050 mol C):(8.10 mol H):(2.700 mol O).
Next, calculate the simplest ratio of moles of elements by dividing the moles of each element by the smallest value in the calculated mole ratio.

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

Divide moles of C by 2.700.

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

Divide moles of H by 2.700.

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

Divide moles of O by 2.700.

The simplest mole ratio is (1.5 mol C):(3 mol H):(1 mol O). Multiply each number in the ratio by the smallest number—in this case 2—that yields a ratio of whole numbers.

$2 \times 1.5 \text{ mol C} = 3 \text{ mol C}$ **Multiply moles of C by 2 to obtain a whole number.**

$2 \times 3 \text{ mol H} = 6 \text{ mol H}$ **Multiply moles of H by 2 to obtain a whole number.**

$2 \times 1 \text{ mol O} = 2 \text{ mol O}$ **Multiply moles of O by 2 to obtain a whole number.**

The simplest whole-number ratio of atoms is (3 atoms C):(6 atoms H):(2 atoms O). Thus, the empirical formula of methyl acetate is **$C_3H_6O_2$**.

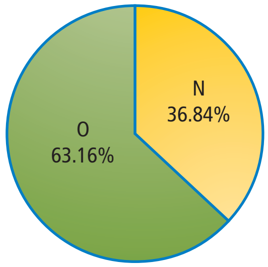
3 Evaluate the Answer

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

PRACTICE Problems

Extra Practice Page 982 and glencoe.com

58. The circle graph at the right gives the percent composition for a blue solid. What is the empirical formula for this solid?



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

60. 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?

61. Challenge Aspirin is the world's most-often used medication. 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.

Molecular Formula

Would it surprise you to learn that 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 ratio in the compound. To identify a new compound, a chemist determine the **molecular formula,** which specifies the actual number of atoms of each element in one molecule or formula unit of the substance. **Figure 10.14** shows an important use of the gas acetylene. It has the same percent composition and the same 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) is 13.02 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.

$$\frac{\text{experimentally determined molar mass of acetylene}}{\text{mass of empirical formula}} = \frac{26.04\ \cancel{\text{g/mol}}}{13.02\ \cancel{\text{g/mol}}} = 2.000$$

Because the molar mass of acetylene is two times the mass represented by the empirical formula, the molecular formula of acetylene must contain twice the number of carbon and hydrogen atoms as represented by the empirical formula.

Figure 10.14 Acetylene is a gas used for welding because of the high-temperature flame produced when it is burned with oxygen.

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.

$$\frac{\text{experimentally determined molar mass of benzene}}{\text{mass of the empirical formula CH}} = \frac{78.12 \text{ g mol}}{13.02 \text{ g mol}} = 6.000$$

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 atoms and hydrogen atoms shown in the empirical formula. You can conclude that the molecular formula for acetylene is $2 \times CH$, or C_2H_2, and the molecular formula for benzene is $6 \times CH$, or C_6H_6.

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

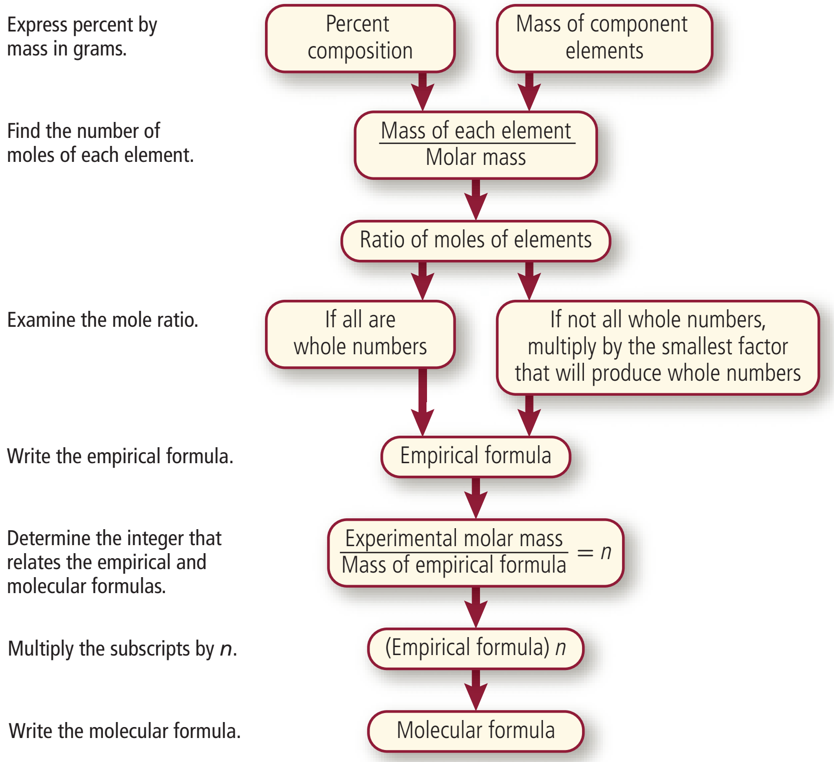
$$\text{molecular formula} = (\text{empirical formula})n$$

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

The steps in determining empirical and molecular formulas from percent composition or mass data are outlined in **Figure 10.15.** 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.

Figure 10.15 Use this flowchart to guide you through the steps in determining the empirical and molecular formulas for compounds.

Describe ***How is the integer* n *related to the empirical and molecular formulas?***



EXAMPLE Problem 10.12

Math Handbook
Ratios
page 964

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.

1 Analyze the Problem

You are given the percent composition. 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

percent by mass C = 40.68% C
percent by mass H = 5.08% H
percent by mass O = 54.24% O
molar mass = 118.1 g/mol succinic acid

Unknown

empirical formula = ?
molecular formula = ?

2 Solve for the Unknown

Use the percents by mass as masses in grams, and convert grams to moles by using a conversion factor—the inverse of molar mass—that relates moles to mass.

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

Substitute mass C = 40.68 g, inverse molar mass C = 1 mol/12.01 g, and solve for moles of C.

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

Substitute mass H = 5.08 g, inverse molar mass H = 1 mol/1.008 g, and solve for moles of H.

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

Substitute mass O = 54.24 g, inverse molar mass O = 1 mol/16.00 g, and solve for moles of O.

The mole ratio in succinic acid is (3.387 mol C):(5.04 mol H):(3.390 mol O).
Next, calculate the simplest ratio of moles of elements by dividing the moles of each element by the smallest value in the calculated mole ratio.

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

Divide moles of C by 3.387.

$$\frac{5.04\ \text{mol H}}{3.387} = 1.49\ \text{mol H} \approx 1.5\ \text{mol H}$$

Divide moles of H by 3.387.

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

Divide moles of O by 3.387.

The simplest mole ratio is 1:1.5:1. Multiply all mole values by 2 to obtain whole numbers.

$2 \times 1\ \text{mol C} = 2\ \text{mol C}$ **Multiply moles of C by 2.**

$2 \times 1.5\ \text{mol H} = 3\ \text{mol H}$ **Multiply moles of H by 2.**

$2 \times 1\ \text{mol O} = 2\ \text{mol O}$ **Multiply moles of O by 2.**

The simplest whole-number mole ratio is 2:3:2. The empirical formula is **$C_2H_3O_2$.**
Calculate the empirical formula mass using the molar mass of each element.

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

Multiply the molar mass of C by the moles of C atoms in the compound.

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

Multiply the molar mass of H by the moles of H atoms in the compound.

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

Multiply the molar mass of O by the moles of O atoms in the compound.

molar mass $C_2H_3O_2$ = (24.02 g + 3.024 g + 32.00 g) = 59.04 g/mol **Total the mass values.**

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

$$n = \frac{\text{molar mass of succinic acid}}{\text{molar mass of } C_2H_3O_2} = \frac{118.1\ \cancel{\text{g/mol}}}{59.04\ \cancel{\text{g/mol}}} = 2.000$$

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

$$2 \times (C_2H_3O_2) = C_4H_6O_4$$

The molecular formula for succinic acid is **$C_4H_6O_4$.**

3 Evaluate the Answer

The calculation of the molar mass from the molecular formula gives the same result as the given, experimentally-determined molar mass.

EXAMPLE Problem 10.13

Math Handbook
Ratios
page 964

Calculating an Empirical Formula from Mass Data The mineral ilmenite is usually mined and processed for titanium, a strong, light, and flexible metal. A sample of ilmenite contains 5.41 g of iron, 4.64 g of titanium, and 4.65 g of 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, then find the smallest whole-number ratio of the moles of the elements.

Known	Unknown
mass of iron = 5.41 g Fe	**empirical formula = ?**
mass of titanium = 4.64 g Ti	
mass of oxygen = 4.65 g O	

2 Solve for the Unknown

Convert each known mass to moles by using a conversion factor—the inverse of molar mass—that relates moles to grams.

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

Substitute mass Fe = 5.41 g, inverse molar mass Fe = 1 mol/55.85 g, and calculate moles of Fe.

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

Substitute mass Ti = 4.64 g, inverse molar mass Ti = 1 mol/47.88 g, and calculate moles of Ti.

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

Substitute mass O = 4.65 g, inverse molar mass O = 1 mol/16.00 g, and calculate moles of O.

The mineral ilmenite has a mole ratio of (0.0969 mol Fe):(0.0969 mol Ti):(0.291 mol O).

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

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

Divide the moles of Fe by 0.0969.

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

Divide the moles of Ti by 0.0969.

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

Divide the moles of O by 0.0969.

Because all the mole values are whole numbers, the simplest whole-number mole ratio is (1 mol Fe):(1 mol Ti):(3 mol O). 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 numbers 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 one-third that of titanium. Thus, a 3:1 ratio of oxygen to titanium is reasonable.

PRACTICE Problems

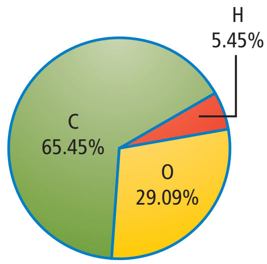
Extra Practice Page 982 and glencoe.com

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

63. 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?

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

65. Challenge Analysis of a chemical used in photographic developing fluid yielded the percent composition data shown in the circle graph to the right. If the chemical's molar mass is 110.0 g/mol, what is its molecular formula?



66. Challenge Analysis of the pain reliever morphine yielded the data shown in the table. Determine the empirical formula of morphine.

Element	Mass (g)
carbon	17.900
hydrogen	1.680
oxygen	4.225
nitrogen	1.228

Section 10.4 Assessment

Section Summary

- The percent by mass of an element in a compound gives the percentage of the compound's total mass due to that element.
- The subscripts in an empirical formula give the smallest whole-number ratio of moles of elements in the compound.
- The molecular formula gives the actual number of atoms of each element in a molecule or formula unit of a substance.
- The molecular formula is a whole-number multiple of the empirical formula.

67. MAIN Idea **Assess** A classmate tells you that experimental data shows a compound's molecular formula to be 2.5 times its empirical formula. Is he correct? Explain.

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

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

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

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

72. Apply The molar mass of a compound is twice that of its empirical formula. How are the compound's molecular and empirical formulas related?

73. Analyze 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?

Chemistry Online Self-Check Quiz glencoe.com

Section 10.5

Objectives

- **Explain** what a hydrate is and relate the name of the hydrate to its composition.
- **Determine** the formula of a hydrate from laboratory data.

Review Vocabulary

crystal lattice: a three-dimensional geometric arrangement of particles

New Vocabulary

hydrate

Formulas of Hydrates

MAIN Idea Hydrates are solid ionic compounds in which water molecules are trapped.

Real-World Reading Link Some products, such as electronic equipment, are boxed with small packets labeled *dessicant*. These packets control moisture by absorbing water. Some contain ionic compounds called hydrates.

Naming Hydrates

Have you ever watched crystals slowly form from a water solution? Sometimes, water molecules adhere to the ions as the solid forms. The water molecules that become part of the crystal are called waters of hydration. Solid ionic compounds 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 10.16** shows the beautiful gemstone known as opal, which is hydrated silicon dioxide (SiO_2). The unusual coloring is the result of water in the mineral.

In the formula of a hydrate, the number of water molecules associated with each formula unit of the compound is written following a dot—for example, $Na_2CO_3{\bullet}10H_2O$. This compound is called sodium carbonate decahydrate. In the word *decahydrate,* the prefix *deca-* means *ten* and the root word *hydrate* refers to *water.* A decahydrate has ten water molecules associated with one formula unit of compound. The mass of water associated with a formula unit is included in molar mass calculations. The number of water molecules associated with hydrates varies widely. Some common hydrates are listed in **Table 10.1.**

Figure 10.16 The presence of water and various mineral impurities accounts for the variety of different-colored opals. Further changes in color occur when opals are allowed to dry out.

Concepts In Motion

Interactive Table Explore naming hydrates at glencoe.com.

Table 10.1 Formulas of Hydrates

Prefix	Molecules H_2O	Formula	Name
Mono-	1	$(NH_4)_2C_2O_4{\bullet}H_2O$	ammonium oxalate monohydrate
Di-	2	$CaCl_2{\bullet}2H_2O$	calcium chloride dihydrate
Tri-	3	$NaC_2H_3O_2{\bullet}3H_2O$	sodium acetate trihydrate
Tetra-	4	$FePO_4{\bullet}4H_2O$	iron(III) phosphate tetrahydrate
Penta-	5	$CuSO_4{\bullet}5H_2O$	copper(II) sulfate pentahydrate
Hexa-	6	$CoCl_2{\bullet}6H_2O$	cobalt(II) chloride hexahydrate
Hepta-	7	$MgSO_4{\bullet}7H_2O$	magnesium sulfate heptahydrate
Octa-	8	$Ba(OH)_2{\bullet}8H_2O$	barium hydroxide octahydrate
Deca-	10	$Na_2CO_3{\bullet}10H_2O$	sodium carbonate decahydrate

The hydrate cobalt(II) chloride hexahydrate is pink.

The hydrate can be heated to drive off the water of hydration.

Anhydrous cobalt(II) chloride is blue.

■ **Figure 10.17** Water of hydration can be removed by heating a hydrate, producing an anhydrous compound that can look very different from its hydrated form.

VOCABULARY

WORD ORIGIN

Anhydrous
comes from the Greek root –*an,* meaning *not* or *without,* and –*hydrous* from the Greek root *hydro* meaning *water*

Analyzing a Hydrate

When a hydrate is heated, water molecules are driven off leaving an anhydrous compound, or one "without water." See **Figure 10.17.** The series of photos show that when pink cobalt(II) chloride hexahydrate is heated, blue anhydrous cobalt(II) chloride is produced.

How can you determine the formula of a hydrate? You must find the number of moles of water associated with 1 mol 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 1 mol 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 2 mol of water is associated with 1 mol of barium chloride. The value of the coefficient x is 2 and the formula of 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 practice in experimentally determining the formula of a hydrate.

Reading Check **Explain** why a dot is used in writing the formula of a hydrate.

EXAMPLE Problem 10.14

Math Handbook
Ratios
page 964

Determining the Formula of 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 of white anhydrous copper sulfate ($CuSO_4$) remains. What is the formula for the hydrate? Name the hydrate.

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 H_2O = 18.02 g/mol H_2O
molar mass $CuSO_4$ = 159.6 g/mol $CuSO_4$

Unknown

formula of hydrate = ?
name of hydrate = ?

2 Solve for the Unknown

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

Subtract the mass of anhydrous $CuSO_4$ from the mass of $CuSO_4 \cdot xH_2O$.

Convert the known masses of H_2O and anhydrous $CuSO_4$ to moles using a conversion factor—the inverse of molar mass—that relates moles and mass.

$$1.59\ \cancel{\text{g } CuSO_4} \times \frac{1 \text{ mol } CuSO_4}{159.6\ \cancel{\text{g } CuSO_4}} = 0.00996 \text{ mol } CuSO_4$$

Substitute mass $CuSO_4$ = 1.59 g, inverse molar mass $CuSO_4$ = 1 mol/159.6 g, and solve.

$$0.91\ \cancel{\text{g } H_2O} \times \frac{1 \text{ mol } H_2O}{18.02\ \cancel{\text{g } H_2O}} = 0.050 \text{ mol } H_2O$$

Substitute mass H_2O = 0.91 g, inverse molar mass H_2O = 1 mol/18.02 g, and solve.

$$x = \frac{\text{moles } H_2O}{\text{moles } CuSO_4}$$

State the ratio of moles of H_2O to moles of $CuSO_4$.

$$x = \frac{0.050 \text{ mol } H_2O}{0.00996 \text{ mol } CuSO_4} \approx \frac{5.0 \text{ mol } H_2O}{1 \text{ mol } CuSO_4} = 5$$

Substitute moles of H_2O = 0.050 mol, moles of $CuSO_4$ = 0.00996 mol. Divide numbers, and cancel units to determine the simplest whole-number ratio.

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

3 Evaluate the Answer

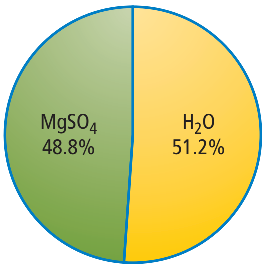
Copper(II) sulfate pentahydrate a common hydrate listed in **Table 10.1.**

PRACTICE Problems

Extra Practice Page 982 and glencoe.com

74. The composition of a hydrate is given in the circle graph shown at the right. What is the formula and name of this hydrate?

75. Challenge An 11.75-g sample of a common hydrate of cobalt(II) chloride is heated. After heating, 0.0712 mol of anhydrous cobalt chloride remains. What is the formula and the name of this hydrate?



■ **Figure 10.18** Calcium chloride, in the bottom of the desiccator, keeps the air inside the desiccator dry. In the chemistry lab, calcium chloride can also be packed into glass tubes called drying tubes. Drying tubes protect reactions from atmospheric moisture, but allow gases produced by reactions to escape.

Uses of Hydrates

Anyhydrous compounds have important applications in the chemistry laboratory. Calcium chloride forms three hydrates—a monohydrate, a dihydrate, and a hexahydrate. As shown in **Figure 10.18,** 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, creating a dry atmosphere in which other substances can be kept dry. Calcium sulfate is often added to solvents such as ethanol and ethyl ether to keep them free of water.

The ability of the anhydrous form of a hydrate to absorb water also has some important commercial applications. Electronic and optical equipment, particularly equipment that is transported overseas by ship, is often packaged with packets of desiccant. Desiccants prevent moisture from interfering with the sensitive electronic circuitry. While some types of desiccant simply absorb moisture, other types bond with moisture from the air and form hydrates.

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 mol of water of hydration. In the process, energy is absorbed by the hydrate. This energy is released when the temperature decreases and the hydrate crystallizes again.

Section 10.5 Assessment

Section Summary

- The formula of a hydrate consists of the formula of 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 1 mol of the compound.
- Anhydrous compounds are formed when hydrates are heated.

76. MAIN Idea **Summarize** the composition of a hydrate.
77. **Name** the compound that has the formula $SrCl_2{\cdot}6H_2O$.
78. **Describe** the experimental procedure for determining the formula of a hydrate. Explain the reason for each step.
79. **Apply** A hydrate contains 0.050 mol of H_2O to every 0.00998 mol of ionic compound. Write a generalized formula of the hydrate.
80. **Calculate** the mass of the water of hydration if a hydrate loses 0.025 mol of H_2O when heated.
81. **Arrange** these hydrates in order of increasing percent water content: $MgSO_4{\cdot}7H_2O$, $Ba(OH)_2{\cdot}8H_2O$, and $CoCl_2{\cdot}6H_2O$.
82. **Apply** Explain how the hydrate in **Figure 10.17** might be used as a means of roughly determining the probability of rain.

Chemistry Online Self-Check Quiz glencoe.com

Everyday Chemistry

History In a Glass of Water

Recall the last glass of water you drank. Although it seems unbelievable, that glass of water almost certainly contained water molecules that were also consumed by Albert Einstein, Joan of Arc, or Confucius! Just how can two glasses of water poured at different times in history contain some of the same molecules? Avogadro's number and molar calculations tell the story.

Oceans and moles The total mass of the water in Earth's oceans and from a variety of other sources is approximately 1.4×10^{24} g. In contrast, an 8-fluid ounce glass of water contains about 2.3×10^{2} g, or 230 g, of water. Using this data, you can calculate the total number of glasses of water available on Earth to drink, and the total number of water molecules contained in those glasses.

You know that one mol of water has a mass of about 18 g. Using dimensional analysis you can convert the grams of water in a glass to moles.

$$\frac{230\ \cancel{\text{g water}}}{\text{glass}} \times \frac{1\ \text{mol water}}{18\ \cancel{\text{g water}}} \approx$$
$$13\ \text{mol water/glass}$$

Thus, one glass of water contains around 13 moles of water. Now convert moles of water to molecules of water by using Avogadro's number.

$$\frac{13\ \cancel{\text{mol water}}}{\text{glass}} \times 6 \times \frac{10^{23}\ \text{molecules water}}{1\ \cancel{\text{mol water}}} \approx$$
$$8 \times 10^{24}\ \text{molecules water/glass}$$

Because you know the total mass of water and the mass of water per glass, you can calculate the total number of glasses of water available for drinking.

$$1.4 \times 10^{24}\ \cancel{\text{g water}} \times \frac{1\ \text{glass}}{230\ \cancel{\text{g water}}} \approx$$
$$6 \times 10^{21}\ \text{glasses}$$

So, there are 8×10^{24} molecules in a single glass of water and there are 6×10^{21} glasses of water on Earth. Comparing these numbers, you can see that there are about 1000 times more molecules in a single glass of water than there are glasses of water on Earth!

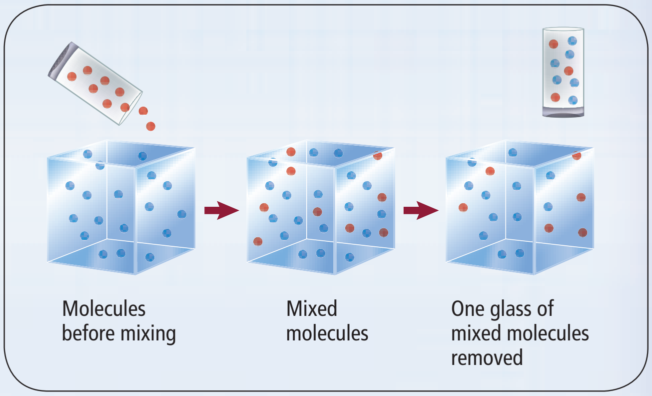


Figure 1 Molecules from the first glass of water (red) are poured back into a container that holds all of Earth's water molecules (blue). A second glass of water taken from the container contains a small number of water molecules that were also in the first glass.

Giant container Suppose all the water on Earth was stored in a single, cube-shaped container. It would be enormous, with sides about 1100 km long! Imagine filling your glass with water from the container. Pour the water back into the container and wait for the water to mix completely. Then refill your glass. Would any of the molecules from the first glass be found in the second glass?

As shown in **Figure 1,** it is likely that the two glasses will share some number of water molecules. Why? Because there are 1000 times more molecules in a glass than there are glasses in the container, on average, the second glass will contain about 1000 molecules that were also in the first glass. This is true for any two glasses.

The power of big numbers Now, consider the amount of water—much more than a single glass—that passed through Einstein, Joan of Arc, or Confucius in their lifetimes. Assuming the molecules of water mixed evenly throughout the entire volume of Earth's water, you can understand how every glass of water must contain some of those same molecules.

WRITING in Chemistry

Estimate The estimating process used in this article is sometimes called a "back-of-the-envelope" calculation. Use this method to estimate the total mass of all of the students in your school. For more on big numbers, visit glencoe.com.

CHEMLAB

DETERMINE THE FORMULA OF A HYDRATE

Chemistry Online

Probeware Alternate CBL instructions can be found at glencoe.com.

Background: In a hydrate, the moles of water to moles of compound ratio is a small whole number. This ratio can be determined by heating the hydrate to remove water.

Question: *How can you determine the moles of water in a mole of a hydrated 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

WARNING: ***Turn off the Bunsen burner when not in use. Crucible, lid, and triangle will be hot and can burn skin. Do not inhale fumes—they are respiratory irritants.***

Procedure

1. Read and complete the lab safety form.
2. Prepare a data table.
3. Measure the mass of the crucible and its lid to the nearest 0.01 g.
4. 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.
5. Record your observations of the hydrate.
6. Place the triangle on the ring of the ring stand. Adjust the ring stand so the triangle will be positioned near the tip of the Bunsen burner's flame. Do not light the Bunsen burner yet.
7. Carefully place the crucible in the triangle with its lid slightly ajar.
8. Begin heating with a low flame, then gradually progress to a stronger flame. Heat for about 10 min, then turn off the burner.
9. Use tongs to carefully remove the crucible from the triangle. Use tongs to place the lid on the crucible. Allow everything to cool.
10. Measure the mass of the crucible, lid, and $MgSO_4$.
11. Record your observations of the anhydrous $MgSO_4$.
12. **Cleanup and Disposal** Discard the anhydrous $MgSO_4$ as directed by your teacher. Return all lab equipment to its proper place and clean your station.

Analyze and Conclude

1. **Calculate** Use your experimental data to calculate the formula for hydrated $MgSO_4$.
2. **Observe and Infer** How do appearances of the hydrated and anhydrous $MgSO_4$ crystals compare? How are they different?
3. **Conclude** Why might the method used not be suitable for determining the water of hydration for all hydrates?
4. **Error Analysis** If the hydrate's formula is $MgSO_4{\cdot}7H_2O$, what is the percent error in your formula for hydrated $MgSO_4$? What are the possible sources for the error? What procedural changes could you make to reduce the error?
5. **Predict** the result of leaving the anhydrous crystals uncovered overnight.

INQUIRY EXTENSION

Design an Experiment to test whether a compound is hydrated or anhydrous.

Study Guide

BIG Idea The mole represents a large number of extremely small particles.

Section 10.1 Measuring Matter

MAIN Idea Chemists use the mole to count atoms, molecules, ions, and formula units.

Vocabulary

- Avogadro's number (p. 321)
- mole (p. 321)

Key Concepts

- The mole is a unit used to count particles of matter indirectly. One mole of a pure substance contains Avogadro's number of particles.
- Representative particles include atoms, ions, molecules, formula units, electrons, and other similar particles.
- One mole of carbon-12 atoms has a mass of exactly 12 g.
- Conversion factors written from Avogadro's relationship can be used to convert between moles and number of representative particles.

Section 10.2 Mass and the Mole

MAIN Idea A mole always contains the same number of particles; however, moles of different substances have different masses.

Vocabulary

- molar mass (p. 326)

Key Concepts

- The mass in grams of 1 mol of any pure substance is called its molar mass.
- The molar mass of an element is numerically equal to its atomic mass.
- 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 to mass. The inverse of molar mass is used to convert from mass to moles.

Section 10.3 Moles of Compounds

MAIN Idea The molar mass of a compound can be calculated from its chemical formula and can be used to convert from mass to moles of that compound.

Key Concepts

- Subscripts in a chemical formula indicate how many moles of each element are present in 1 mol of the compound.
- The molar mass of a compound is calculated from the molar masses of all of the elements in the compound.
- Conversion factors based on a compound's molar mass are used to convert between moles and mass of a compound.

Section 10.4 Empirical and Molecular Formulas

MAIN Idea A molecular formula of a compound is a whole-number multiple of its empirical formula.

Vocabulary

- empirical formula (p. 344)
- molecular formula (p. 346)
- percent composition (p. 342)

Key Concepts

- The percent by mass of an element in a compound gives the percentage of the compound's total mass due to that element.
- The subscripts in an empirical formula give the smallest whole-number ratio of moles of elements in the compound.
- The molecular formula gives the actual number of atoms of each element in a molecule or formula unit of a substance.
- The molecular formula is a whole-number multiple of the empirical formula.

Section 10.5 Formulas of Hydrates

MAIN Idea Hydrates are solid ionic compounds in which water molecules are trapped.

Vocabulary

- hydrate (p. 351)

Key Concepts

- The formula of a hydrate consists of the formula of the ionic compound and the number of water molecules associated with one formula unit.
- The name of a hydrate consists of the compound name and the word *hydrate* with a prefix indicating the number of water molecules in 1 mol of the compound.
- Anhydrous compounds are formed when hydrates are heated.

CHAPTER 10 Assessment

Section 10.1

Mastering Concepts

83. What is the numerical value of Avogadro's number?

84. How many atoms of potassium does 1 mol of potassium contain?

85. Compare a mole of Ag-108 and a mole of Pt-195 using atoms, protons, electrons, and neutrons.

86. Why is the mole an important unit to chemists?

87. **Currency** Examine the information in **Table 10.2** and explain how rolls used to count pennies and dimes are similar to moles.

Table 10.2 Rolled-Coin Values

Coin	Value of a Roll of Coins
Penny	\$0.50
Dime	\$5.00

88. Explain how Avogadro's number is used as a conversion factor.

89. **Conversion** Design a flowchart that could be used to help convert particles to moles or moles to particles.

Mastering Problems

90. Determine the number of representative particles in each substance.
a. 0.250 mol of silver
b. 8.56×10^{-3} mol of sodium chloride
c. 35.3 mol of carbon dioxide
d. 0.425 mol of nitrogen (N_2)

91. Determine the number of representative particles in each substance.
a. 4.45 mol of $C_6H_{12}O_6$
b. 0.250 mol of KNO_3
c. 2.24 mol of H_2
d. 9.56 mol of Zn

92. How many molecules are contained in each compound?
a. 1.35 mol of carbon disulfide (CS_2)
b. 0.254 mol of diarsenic trioxide (As_2O_3)
c. 1.25 mol of water
d. 150.0 mol of HCl

93. Determine the number of moles in each substance.
a. 3.25×10^{20} atoms of lead
b. 4.96×10^{24} molecules of glucose
c. 1.56×10^{23} formula units of sodium hydroxide
d. 1.25×10^{25} copper(II) ions

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

95. How many moles contain the given quantity?
a. 1.25×10^{15} molecules of carbon dioxide
b. 3.59×10^{21} formula units of sodium nitrate
c. 2.89×10^{27} formula units of calcium carbonate

96. **RDA of Selenium** The recommended daily allowance (RDA) of selenium in your diet is 8.87×10^{-4} mol. How many atoms of selenium is this?

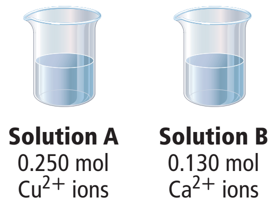


Figure 10.19

97. The two solutions shown in **Figure 10.19** are mixed. What is the total number of metal ions in the mixture?

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

99. **Snowflakes** A snowflake contains 1.9×10^{18} molecules of water. How many moles of water does it contain?

100. 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 for 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?

101. **Chlorophyll** The green color of leaves is due to the presence of chlorophyll, $C_{55}H_{72}O_5N_4Mg$. A fresh leaf was found to have 1.5×10^{-5} mol of chlorophyll per cm^2. How many chlorophyll molecules are in 1 cm^2?

Section 10.2

Mastering Concepts

102. Explain the difference between atomic mass (amu) and molar mass (g).

103. Which contains more atoms, a mole of silver atoms or a mole of gold atoms? Explain your answer.

104. Which has more mass, a mole of potassium or a mole of sodium? Explain your answer.

105. Explain how you would convert from number of atoms of a specific element to its mass.

106. Discuss the relationships that exist between the mole, molar mass, and Avogadro's number.

107. **Barbed Wire** Barbed wire is often made of steel, which is primarily iron, and coated with zinc. Compare the number of particles and the mass of 1 mol of each.

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Mastering Problems

108. Calculate the mass of each element.
a. 5.22 mol of He
b. 0.0455 mol of Ni
c. 2.22 mol of Ti
d. 0.00566 mol of Ge

109. Perform the following conversions.
a. 3.50 mol of Li to g of Li
b. 7.65 g of Co to mol of Co
c. 5.62 g of Kr to mol of Kr
d. 0.0550 mol of As to g of As

110. Determine the mass in grams of each element.
a. 1.33×10^{22} mol of Sb
b. 4.75×10^{14} mol of Pt
c. 1.22×10^{23} mol of Ag
d. 9.85×10^{24} mol of Cr

111. Complete **Table 10.3.**

Table 10.3 Mass, Mole, and Particle Data

Mass	Moles	Particles
	3.65 mol Mg	
29.54 g Cr		
		3.54×10^{25} atoms P
	0.568 mol As	

112. Convert each 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

113. Calculate the number of atoms in each element.
a. 25.8 g of Hg
b. 0.0340 g of Zn
c. 150 g of Ar
d. 0.124 g of Mg

114. Arrange from least to most in moles: 3.00×10^{24} atoms Ne, 4.25 mole Ar, 2.69×10^{24} atoms Xe, 65.96 g Kr.

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

116. A sample of a compound contains 3.86 g of sulfur and 4.08 g of vanadium. How many atoms of sulfur and vanadium does the compound contain?

117. Which has more atoms, 10.0 g of C or 10.0 g of Ca? How many atoms does each have?

118. Which has more atoms, 10.0 mol of C or 10.0 mol of Ca? How many atoms does each have?

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

120. Respiration Air contains several gases. When resting, every breath you take contains approximately 0.600 g of air. If argon makes up 0.934% of the air, calculate the number of argon atoms inhaled with each breath.

Section 10.3

Mastering Concepts

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

122. 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?

123. Explain how you determine the molar mass of a compound.

124. Insect Repellent Many insect repellents use DEET as the active ingredient. DEET was patented in 1946 and is effective against many biting insects. What must you know to determine the molar mass of DEET?

125. Why can molar mass be used as a conversion factor?

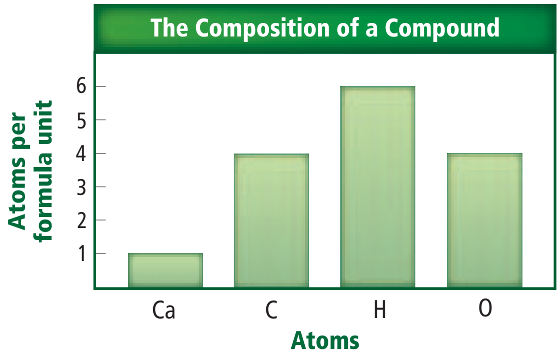
126. List three conversion factors used in molar conversions.

127. Which of these 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.

Mastering Problems

128. How many moles of oxygen atoms are contained in each compound?
a. 2.50 mol of $KMnO_4$
b. 45.9 mol of CO_2
c. 1.25×10^{-2} mol of $CuSO_4 \cdot 5H_2O$

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



■ **Figure 10.20**

130. The graph in **Figure 10.20** shows the numbers of atoms of each element in a compound. What is the compound's formula? What is its molar mass?

131. Determine the molar mass of each compound.
a. nitric acid (HNO_3)
b. ammonium nitrate (NH_4NO_3)
c. zinc oxide (ZnO)
d. cobalt chloride ($CoCl_2$)

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

133. How many moles are in 100.0 g of each compound?
a. dinitrogen oxide (N_2O)
b. methanol (CH_3OH)

134. What is the mass of each compound?
a. 4.50×10^{-2} mol of $CuCl_2$
b. 1.25×10^{2} mol of $Ca(OH)_2$

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

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

137. What is the mass of a mole of electrons if one electron has a mass of 9.11×10^{-28} g?

138. How many moles of ions are in each compound?
a. 0.0200 g of $AgNO_3$
b. 0.100 mol of K_2CrO_4
c. 0.500 g of $Ba(OH)_2$
d. 1.00×10^{-9} mol of Na_2CO_3

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

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

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

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

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

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

145. Cooking A common cooking vinegar is 5.0% acetic acid (CH_3COOH). How many molecules of acetic acid are present in 25.0 g of vinegar?

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

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

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

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

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

151. Espresso There is 1.00×10^{2} mg of caffeine in a shot of espresso. The chemical formula of caffeine is $C_8H_{10}N_4O_2$. Determine the moles of each element present in the caffeine in one shot of espresso.

152. The density of lead (Pb) is 11.3 g/cm^3. Calculate the volume of 1 mol of Pb.

Section 10.4

Mastering Concepts

153. Explain what is meant by percent composition.

154. What information must a chemist obtain in order to determine the empirical formula of an unknown compound?

155. What information must a chemist have to determine the molecular formula for a compound?

156. What is the difference between an empirical formula and a molecular formula? Provide an example.

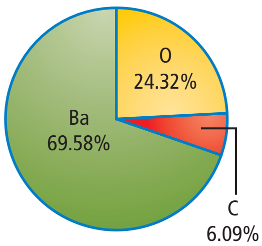
157. When can the empirical formula be the same as the molecular formula?

158. Antibacterial Soap Triclosan is an antibacterial agent included in detergents, dish soaps, laundry soaps, deodorants, cosmetics, lotions, creams, toothpastes, and mouthwashes. The chemical formula for triclosan is $C_{12}H_7Cl_3O_2$. What information did the chemist need to determine this formula?

159. Which of the following formulas—NO, N_2O, NO_2, N_2O_4, and N_2O_5—represent the empirical and molecular formulas of the same compound? Explain your answer.

160. Do all pure samples of a given compound have the same percent composition? Explain.

Mastering Problems

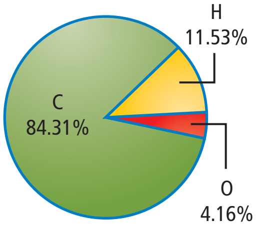


■ **Figure 10.21**

161. The circle graph in **Figure 10.21** shows the percent composition of a compound containing barium, carbon, and oxygen. What is the empirical formula of this compound?

162. Iron Three naturally occurring iron compounds are pyrite (FeS_2), hematite (Fe_2O_3), and siderite ($FeCO_3$). Which contains the greatest percentage of iron?

163. Express the composition of each compound as the mass percent of its elements (percent composition).
a. sucrose ($C_{12}H_{22}O_{11}$) **c.** magnetite (Fe_3O_4)
b. aluminum sulfate ($Al_2(SO_4)_3$)



Molar mass = 384 g/mol

■ **Figure 10.22**

164. Vitamin D_3 Your body's ability to absorb calcium is aided by vitamin D_3. Chemical analysis of vitamin D_3 yields the data shown in **Figure 10.22.** What are the empirical and molecular formulas for vitamin D_3?

165. When a 35.07-g sample of phosphorus reacts with oxygen, a 71.00-g sample of phosphorus oxide is formed. What is the percent composition of the compound? What is the empirical formula for this compound?

166. Cholesterol Heart disease is linked to high blood cholesterol levels. What is the percent composition of the elements in a molecule of cholesterol ($C_{27}H_{45}OH$)?

167. Determine the empirical formula for each compound.
a. ethylene (C_2H_4)
b. ascorbic acid ($C_6H_8O_6$)
c. naphthalene ($C_{10}H_8$)

168. Caffeine The stimulant effect of coffee is due to caffeine, $C_8H_{10}N_4O_2$. Calculate the molar mass of caffeine. Determine its percent composition.

169. Which titanium-containing mineral, rutile (TiO_2) or ilmenite ($FeTiO_3$), has the larger percentage of titanium?

170. Vitamin E Many plants contain vitamin E ($C_{29}H_{50}O_2$), a substance that some think slows the aging process in humans. What is the percent composition of vitamin E?

171. Artificial Sweetener Determine the percent composition of aspartame ($C_{14}H_{18}N_2O_5$), an artificial sweetener.

172. MSG Monosodium glutamate, known as 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 its empirical formula?

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

174. Patina The Statue of Liberty has turned green because of the formation of a patina. Two copper compounds, $Cu_3(OH)_4SO_4$ and $Cu_4(OH)_6SO_4$, form this patina. Find the mass percentage of copper in each compound.

Section 10.5

Mastering Concepts

175. What is a hydrated compound? Use an example to illustrate your answer.

176. Explain how hydrates are named.

177. Desiccants Why are certain electronic devices transported with desiccants?

178. In a laboratory setting, how would you determine if a compound was a hydrate?

179. Write the formula for the following hydrates.
a. nickel(II) chloride hexahydrate
b. cobalt(II) chloride hexahydrate
c. magnesium carbonate pentahydrate
d. sodium sulfate decahydrate

Mastering Problems

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

181. Table 10.4 shows data from an experiment to determine the formulas of hydrated barium chloride. Determine the formula for the hydrate and its name.

Table 10.4 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	

182. Chromium(III) nitrate forms a hydrate that is 40.50% water by mass. What is its chemical formula?

183. Determine the percent composition of $MgCO_3{\cdot}5H_2O$ and draw a pie graph to represent the hydrate.

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

185. Gypsum is hydrated calcium sulfate. A 4.89-g sample of this hydrate was heated. After the water was removed, 3.87 g anhydrous calcium sulfate remained. Determine the formula for this hydrate and name the compound.

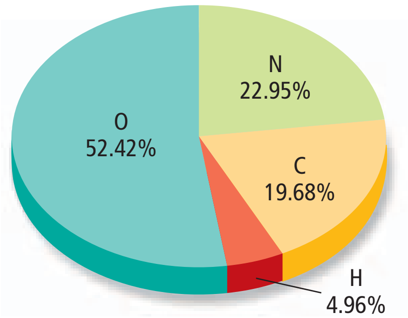
186. A 1.628-g sample of a hydrate of magnesium iodide is heated until its mass is reduced to 1.072 g and all water has been removed. What is the formula of the hydrate?

187. Borax 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.

Chapter 10 Assessment

Mixed Review

188. Rank samples A–D from least number of atoms to greatest number of atoms. A: 1.0 mol of H_2; B: 0.75 mol of H_2O; C: 1.5 mol of NaCl; D: 0.50 mol of Ag_2S



■ **Figure 10.23**

189. The graph in **Figure 10.23** 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?

190. How many grams of $CoCl_2 \cdot 6H_2O$ must you measure out in a container to have exactly Avogadro's number of particles?

191. One atom of an unknown element has a mass of 6.66×10^{-23} g. What is the identity of this element?

192. Skunks Analysis of skunk spray yields a molecule with 44.77% C, 7.46% H and 47.76% S. What is the chemical formula for this molecule found in the spray from skunks that scientists think is partly responsible for the strong odor?

193. How many moles are present in 1.00 g of each compound?

a. L-tryptophan ($C_{11}H_{12}N_2O_2$), an essential amino acid

b. magnesium sulfate heptahydrate, also known as Epsom salts

c. propane (C_5H_8), a fuel

194. A compound contains 6.0 g of carbon and 1.0 g of hydrogen, and has a molar mass of 42.0 g/mol. What are the compound's percent composition, empirical formula, and molecular formula?

195. Which of these compounds has the greatest percent of oxygen by mass: TiO_2, Fe_2O_3, or Al_2O_3?

196. Mothballs Naphthalene, commonly found in mothballs, is composed of 93.7% carbon and 6.3% hydrogen. The molar mass of naphthalene is 128 g/mol. Determine the empirical and molecular formulas for naphthalene.

197. Which of these molecular formulas are also empirical formulas: ethyl ether ($C_4H_{10}O$), aspirin ($C_9H_8O_4$), butyl dichloride ($C_4H_8O_2$), glucose ($C_6H_{12}O_6$)?

Think Critically

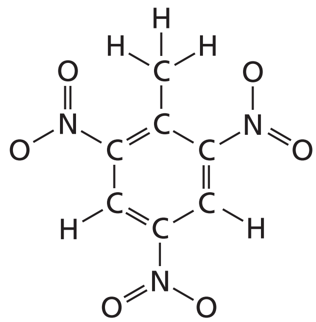
198. Apply Concepts A mining company has two possible sources of copper: chalcopyrite ($CuFeS_2$) and chalcocite (Cu_2S). If the mining conditions and the extraction of copper from the ore were identical for each of the ores, which ore would yield the greater quantity of copper? Explain your answer.

199. Analyze and Conclude On a field trip, students collected rock samples. Analysis of the rocks revealed that two of the rock samples contained lead and sulfur. **Table 10.5** shows the percent lead and sulfur in each of the rocks. Determine the molecular formula of each rock. What can the students conclude about the rock samples?

Table 10.5 Lead and Sulfur Content

Rock Sample	% Lead	% Sulfur
1	86.6 %	13.4%
2	76.4%	23.6%

200. Graph A YAG, or yttrium aluminum garnet ($Y_3Al_5O_{12}$), is a synthetic gemstone which has no counterpart in nature. Design a bar graph to indicate the moles of each element present in a 5.67 carat yttrium aluminum garnet. (1 carat = 0.20 g)



■ **Figure 10.24**

201. Assess The structure of the TNT molecule is shown in **Figure 10.24.** Critique the statement "Trinitrotoluene, TNT, contains 21 atoms per mole." What is correct about the statement and what is incorrect? Rewrite the statement.

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

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

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Challenge Problem

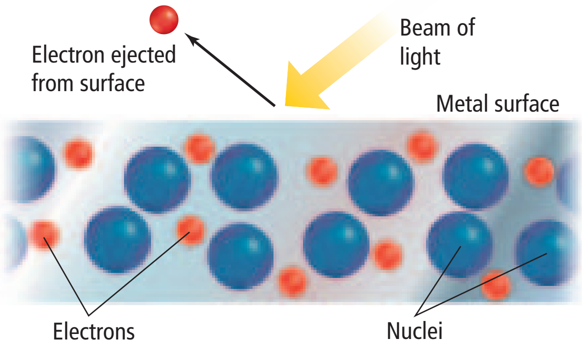
204. Two different compounds are composed of Elements X and Y. The formulas of the compounds are X_2Y_3 and XY. A 0.25 mol sample of XY has a mass of 17.96 g, and a 0.25 mol sample of X_2Y_3 has a mass of 39.92 g.
a. What are the atomic masses of elements X and Y?
b. What are the formulas for the compounds?

Cumulative Review

205. Express each answer 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

206. Making Candy A recipe for pralines calls for the candy mixture to be heated until it reaches the "soft ball" stage, at about 236°F. Can a Celsius thermometer with a range of −10 to 110°C be used to determine when the "soft ball" stage is reached? *(Chapter 2)*

207. Contrast atomic number and mass number. Compare these numbers for isotopes of an element. *(Chapter 4)*



■ **Figure 10.25**

208. Describe the phenomenon in **Figure 10.25.** Explain why the electrons are not bound to the nuclei. *(Chapter 5)*

209. Given the elements Ar, Cs, Br, and Ra, identify those that form positive ions. Explain your answer. *(Chapter 7)*

210. Write the formula and name the compound formed when each pair of elements combine. *(Chapter 7)*
a. barium and chlorine
b. aluminum and selenium
c. calcium and phosphorus

211. Write balanced equations for each reaction. *(Chapter 9)*
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.

Additional Assessment

WRITING in Chemistry

212. Natural Gas Natural gas hydrates are chemical compounds known as clathrate hydrates. Research natural gas hydrates and prepare an educational pamphlet for consumers. The pamphlet should discuss the composition and structure of the compounds, the location of the hydrates, their importance to consumers, and the environmental impact of using the hydrates.

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

214. Luminol Crime-scene investigators use luminol to visualize blood residue. Research luminol and determine its chemical formula and percent composition.

DBQ Document-Based Questions

Space Shuttle Propellants *At liftoff, the orbiter and an external fuel tank carry 3,164,445 L of the liquid propellants hydrogen, oxygen, hydrazine, monomethylhydrazine, and dinitrogen tetroxide. Their total mass is 727,233 kg. Data for the propellants carried at liftoff are given in* **Table 10.6.**

Data obtained from: "Space Shuttle Use of Propellants and Fluids." September 2001. *NASA Fact Sheet.*

Table 10.6 Space Shuttle Liquid Propellants

Propellants	Molecular Formula	Mass (kg)	Moles	Molecules
Hydrogen	H_2		5.14×10^7	
Oxygen	O_2			1.16×10^{31}
Hydrazine		493		
Monomethyl-hydrazine	CH_3NHNH_2	4909		
Dinitrogen tetroxide	N_2O_4		8.64×10^4	

215. Hydrazine contains 87.45% nitrogen and 12.55% hydrogen, and has a molar mass of 32.04 g/mol. Determine hydrazine's molecular formula. Record the molecular formula in **Table 10.6.**

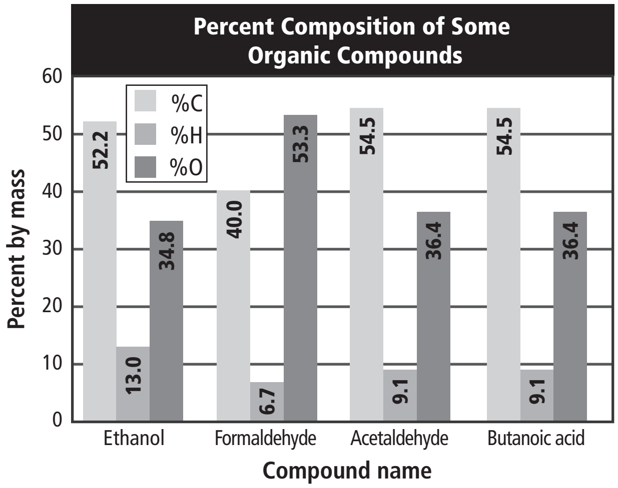
216. Complete **Table 10.6** by calculating the number of moles, mass in kilograms, or molecules for each propellant. Give all answers to three significant figures.

Cumulative

Standardized Test Practice

Multiple Choice

Use the graph below to answer Questions 1 to 4.



1. Acetaldehyde and butanoic acid must have the same
A. molecular formula.
B. empirical formula.
C. molar mass.
D. chemical properties.

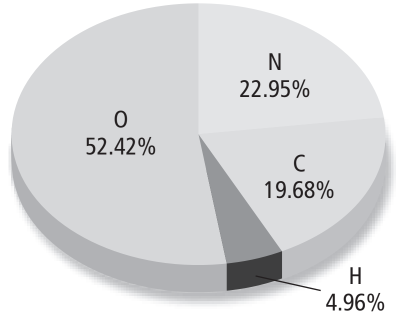
2. If the molar mass of butanoic acid is 88.1 g/mol, what is its molecular formula?
A. $C_3H_4O_3$
B. C_2H_4O
C. $C_5H_{12}O_1$
D. $C_4H_8O_2$

3. What is the empirical formula of ethanol?
A. C_4HO_3
B. $C_2H_6O_2$
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 mol of formaldehyde?
A. 30.00 g
B. 60.06 g
C. 182.0 g
D. 200.0 g

5. Which does NOT describe a mole?
A. a unit used to count particles directly
B. Avogadro's number of molecules of a compound
C. the number of atoms in exactly 12 g of pure C-12
D. the SI unit for the amount of a substance

Use the graph below to answer Question 6.



6. What is the empirical formula for this compound?
A. $C_6H_2N_6O_3$
B. $C_4HN_5O_{10}$
C. CH_3NO_2
D. CH_5NO_3

7. Which is NOT true of molecular compounds?
A. Triple bonds are stronger than single bonds.
B. Electrons are shared in covalent bonds.
C. All atoms have eight valence electrons when they are chemically stable.
D. Lewis structures show the arrangements of electrons in covalent molecules.

8. Which type of reaction is shown below?

$$2HI + (NH_4)_2S \longrightarrow H_2S + 2NH_4I$$

A. synthesis
B. decomposition
C. single replacement
D. double replacement

9. 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}

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

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Short Answer

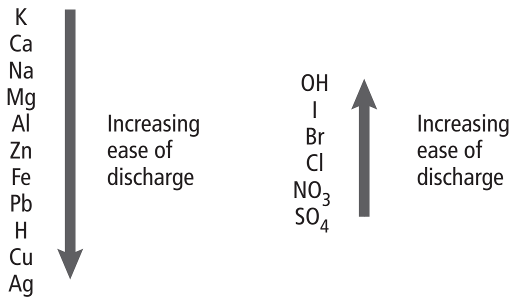
Use the table below to answer Question 11.

Charges of Some Ions	
Ion	**Formula**
Sulfide	S^{-2}
Sulfite	SO_3^{-2}
Sulfate	SO_4^{-2}
Thiosulfate	$S_2O_3^{-2}$
Copper(I)	Cu^{+}
Copper(II)	Cu^{2+}

11. How many possible compounds can be made that contain copper, sulfur, and oxygen? Write their names and formulas.

Extended Response

Use the figure below to answer Question 12.



You have been asked to identify a sample of a metal. It is known to be either zinc, lead, or lithium. You have aqueous solutions of KCl, $AlCl_3$, $FeCl_3$, and $CuCl_2$ available.

12. Explain how you would use these solutions to identify what metal your sample is made of.

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13. It takes 2 iron atoms and 6 chlorine atoms to make 2 iron(III) chloride particles. How many chlorine atoms are required to make 18 iron(III) chloride particles?

A. 9
B. 18
C. 27
D. 54
E. 72

14. 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
E. 524 g/mol

15. Which is not a correct formula for an ionic compound?

A. $CaCl_2$
B. Na_2SO_4
C. Al_3S_2
D. $Mg(NO_3)_2$
E. NaCl

Use the table below to answer question 16.

Percent composition of selected hydrocarbons			
Compound	**%C**	**%H**	**%O**
$C_4H_{10}O$	64.81	13.60	21.59
$C_6H_{12}O_4$	48.64	8.108	43.24
$C_7H_{16}O_3$	56.76	10.81	32.43
$C_5H_8O_5$	40.54	5.405	54.05

16. A 25.0-g sample of an unknown hydrocarbon is composed of 12.16 g carbon, 2.027 g hydrogen, and 10.81 g oxygen. If its molecular weight is 148 g/mol, what is the molecular formula for this compound?

A. $C_4H_{10}O$
B. $C_6H_{12}O_4$
C. $C_7H_{16}O_3$
D. $C_5H_8O_5$
E. $C_8H_5O_5$

NEED EXTRA HELP?																
If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16
Review Section . . .	10.4	10.4	10.4	10.4	10.4	10.2	8.3	9.2	10.1	10.2	7.3	9.2	9.1	10.3	7.3	10.4