10 Chemical Quantities

INSIDE:

- **10.1** The Mole: A Measurement of Matter
- **10.2** Mole–Mass and Mole–Volume Relationships
- **10.3** Percent Composition and Chemical Formulas

PearsonChem.com

When you shop at the grocery store or farmers' market, you usually buy blueberries by the pint, not by the berry. Similarly, chemists use a unit called the mole to count atoms and molecules.

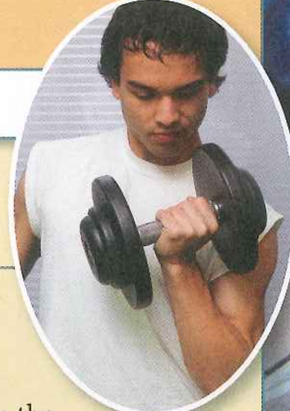
BIGIDEA

THE MOLE AND QUANTIFYING MATTER

Essential Questions:

1. *Why is the mole an important measurement in chemistry?*
2. *How can the molecular formula of a compound be determined experimentally?*

CHEMYSTERY



A Formula for Cheating

Anabolic steroids are compounds that are developed to increase muscle size and strength. Stories are often in the news about professional athletes, such as baseball players, cyclists, and track stars, who have used steroids to enhance their performance.

More than 100 different types of anabolic steroids have been developed, and each of these substances is illegal in the United States without a prescription. Steroids have also been banned by many sports organizations because of their dangerous side effects and because they give the user an unfair advantage. Therefore, athletes are often tested for steroid use. So, how can the presence of steroids in the body be detected?

► Connect to the **BIG**IDEA As you read about the mole and chemical quantities, think about how the molar mass and molecular formula of a compound can be determined and used to identify the presence of steroids in the body.

NATIONAL SCIENCE EDUCATION STANDARDS

A-1, A-2, B-2, E-2, G-1, G-3

10.1 The Mole: A Measurement of Matter

CHEMISTRY & YOU

Q: *How can you quantify the amount of sand in a sand sculpture?* Have you ever gone to the beach and created a castle or sculpture out of sand? You could measure the amount of sand in a sculpture by counting the grains of sand. Is there an easier way to measure the amount of sand? Chemists measure the amount of a substance using a unit called the mole.

Key Questions

How can you convert among the count, mass, and volume of something?

How do chemists count the number of atoms, molecules, or formula units in a substance?

How do you determine the molar mass of an element and of a compound?

Vocabulary

- mole
- Avogadro's number
- representative particle
- molar mass

Measuring Matter

How can you convert among the count, mass, and volume of something?

Chemistry is a quantitative science. In your study of chemistry, you will analyze the composition of samples of matter and perform chemical calculations that relate quantities of the reactants in a chemical reaction to quantities of the products. To solve these and other problems, you will have to be able to measure the amount of matter you have.

One way to measure matter is to count how many of something you have. For example, you can count the mp3s in your collection. Another way to measure matter is to determine its mass. You can buy apples by the kilogram or pound, as shown in Figure 10.1. You can also measure matter by volume. For instance, people buy gasoline by the liter or the gallon.

Some of the units used for measuring indicate a specific number of items. For example, a pair always means two. A pair of shoes is two shoes, and a pair of aces is two aces. Similarly, a dozen always means 12. A dozen eggs is 12 eggs, and a dozen pens is 12 pens.

Apples can be measured in three different ways. At a fruit stand, they are often sold by the count. In a supermarket, you usually buy apples by weight or mass. At an orchard, you can buy apples by volume. Each of these different ways to measure apples can be equated to a dozen apples.

$$\text{By count: 1 dozen apples} = \text{12 apples}$$

For average-sized apples, the following approximations can be used.

$$\text{By mass: 1 dozen apples} = \text{2.0 kg apples}$$

$$\text{By volume: 1 dozen apples} = \text{0.20 bushel apples}$$

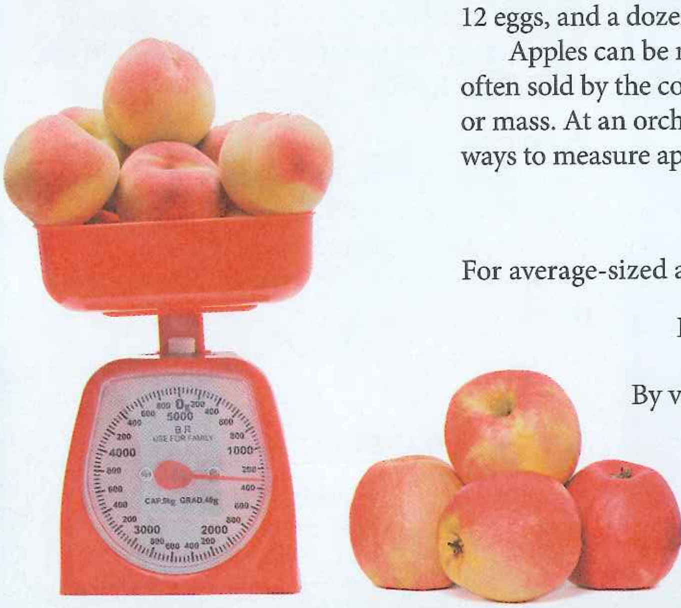


Figure 10.1 Measuring by Mass
A dozen apples has a mass of about 2.0 kg.

Knowing how the count, mass, and volume of an item relate to a common unit allows you to convert among these units. For example, based on the unit relationships given on the previous page, you could calculate the mass of a bushel of apples or the mass of 90 average-sized apples using conversion factors such as the following:

$$\frac{1\ \text{dozen apples}}{12\ \text{apples}} \qquad \frac{2.0\ \text{kg apples}}{1\ \text{dozen apples}} \qquad \frac{1\ \text{dozen apples}}{0.20\ \text{bushel apples}}$$

Sample Problem 10.1

Finding Mass From a Count

What is the mass of 90 average-sized apples if 1 dozen of the apples has a mass of 2.0 kg?

KNOWNS

number of apples = 90 apples
12 apples = 1 dozen apples
1 dozen apples = 2.0 kg apples

UNKNOWN

mass of 90 apples = ? kg

1 Analyze List the knowns and the unknown. Use dimensional analysis to convert the number of apples to the mass of apples.

2 Calculate Solve for the unknown.

First, identify the sequence of conversions needed to perform the calculation.

number of apples ⟶ dozens of apples ⟶ mass of apples

Write the conversion factor to convert from number of apples to dozens of apples.

$$\frac{1\ \text{dozen apples}}{12\ \text{apples}}$$

Write the conversion factor to convert from dozens of apples to mass of apples.

$$\frac{2.0\ \text{kg apples}}{1\ \text{dozen apples}}$$

The units apples and dozen apples cancel, so the answer has the unit kg.

Multiply the number of apples by these two conversion factors to get the answer in kilograms.

$$90\ \cancel{\text{apples}} \times \frac{1\ \cancel{\text{dozen apples}}}{12\ \cancel{\text{apples}}} \times \frac{2.0\ \text{kg apples}}{1\ \cancel{\text{dozen apples}}} = 15\ \text{kg apples}$$

3 Evaluate Does the result make sense? A dozen apples has a mass of 2.0 kg, and 90 apples is less than 10 dozen apples, so the mass should be less than 20 kg of apples (10 dozen × 2.0 kg/dozen).

1. If 0.20 bushel is 1 dozen apples and a dozen apples has a mass of 2.0 kg, what is the mass of 0.50 bushel of apples?

2. Assume 2.0 kg of apples is 1 dozen and that each apple has 8 seeds. How many apple seeds are in 14 kg of apples?

In Problem 1, the desired conversion is bushels of apples ⟶ dozens of apples ⟶ mass of apples.

In Problem 2, the desired conversion is mass of apples ⟶ dozens of apples ⟶ number of apples ⟶ number of seeds.

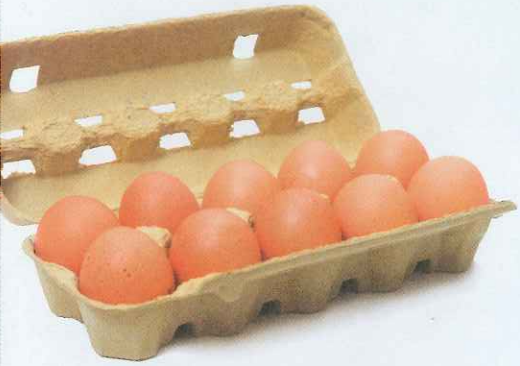


Figure 10.2 Grouping Objects Words other than *mole* are used to describe a number of something—for example, a *dozen* eggs is 12 eggs.

What Is a Mole?

How do chemists count the number of atoms, molecules, or formula units in a substance?

Counting objects as big as apples is a reasonable way to measure how much of the object you have. Picture trying to count the grains of sand in a sand sculpture. It would be an endless job. Recall that matter is composed of atoms, molecules, and ions. These particles are much, much smaller than grains of sand, and an extremely large number of them are in a small sample of a substance. Obviously, counting particles one by one is not practical. However, think about counting eggs. It's easier when the eggs are grouped into dozens, as shown in Figure 10.2. A dozen is a specified number (12) of things.

Counting With Moles Chemists also use a unit that is a specified number of particles. The unit is called the mole. A **mole** (mol) of a substance is 6.02×10^{23} representative particles of that substance and is the SI unit for measuring the amount of a substance. The number of representative particles in a mole, 6.02×10^{23}, is called **Avogadro's number.** It was named in honor of the Italian scientist Amedeo Avogadro di Quaregna (1776–1856), who helped clarify the difference between atoms and molecules.

The term **representative particle** refers to the species present in a substance, usually atoms, molecules, or formula units. The representative particle of most elements is the atom. Iron is composed of iron atoms. Helium is composed of helium atoms. Seven elements, however, normally exist as diatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2). The representative particle of these elements and of all molecular compounds is the molecule. The molecular compounds water (H_2O) and sulfur dioxide (SO_2) are composed of H_2O and SO_2 molecules, respectively. For ionic compounds, such as calcium chloride, the representative particle is the formula unit $CaCl_2$. **The mole allows chemists to count the number of representative particles in a substance.** A mole of any substance contains Avogadro's number of representative particles, or 6.02×10^{23} representative particles. Table 10.1 summarizes the relationship between representative particles and moles of substances.

CHEMISTRY & YOU

Q: *What are the different ways you can measure the amount of sand in a sand sculpture?*

Table 10.1

Representative Particles and Moles

Substance	Representative particle	Chemical formula	Representative particles in 1.00 mol
Copper	Atom	Cu	6.02×10^{23}
Atomic nitrogen	Atom	N	6.02×10^{23}
Nitrogen gas	Molecule	N_2	6.02×10^{23}
Water	Molecule	H_2O	6.02×10^{23}
Sucrose	Molecule	$C_{12}H_{22}O_{11}$	6.02×10^{23}
Calcium ion	Ion	Ca^{2+}	6.02×10^{23}
Calcium fluoride	Formula unit	CaF_2	6.02×10^{23}

Converting Between Number of Particles and Moles The relationship, 1 mol = 6.02×10^{23} representative particles, is the basis for the following conversion factors that you can use to convert number of representative particles to moles and moles to number of representative particles.

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

Sample Problem 10.2

Converting Number of Atoms to Moles

Magnesium is a light metal used in the manufacture of aircraft, automobile wheels, and tools. How many moles of magnesium is 1.25×10^{23} atoms of magnesium?

KNOWN

number of atoms = 1.25×10^{23} atoms Mg

UNKNOWN

moles = ? mol Mg

1 Analyze **List the known and the unknown.** The desired conversion is atoms $\longrightarrow$ moles.

2 Calculate **Solve for the unknown.**

First, state the relationship between moles and number of representative particles.

$$1\ \text{mol Mg} = 6.02 \times 10^{23}\ \text{atoms Mg}$$

Write the conversion factors you get based on this relationship.

$$\frac{1\ \text{mol Mg}}{6.02 \times 10^{23}\ \text{atoms Mg}} \quad \text{and} \quad \frac{6.02 \times 10^{23}\ \text{atoms Mg}}{1\ \text{mol Mg}}$$

Identify the conversion factor needed to convert from atoms to moles.

$$\frac{1\ \text{mol Mg}}{6.02 \times 10^{23}\ \text{atoms Mg}}$$

Multiply the number of atoms of Mg by the conversion factor.

$$1.25 \times 10^{23}\ \cancel{\text{atoms Mg}} \times \frac{1\ \text{mol Mg}}{6.02 \times 10^{23}\ \cancel{\text{atoms Mg}}}$$

$$= 0.208\ \text{mol Mg}$$

3 Evaluate **Does the result make sense?** The given number of atoms (1.25×10^{23}) is less than one fourth of Avogadro's number (6.02×10^{23}), so the answer should be less than one fourth (0.25) mol of atoms. The answer should have three significant figures.

3. How many moles is 2.80×10^{24} atoms of silicon?

4. How many moles is 2.17×10^{23} representative particles of bromine?

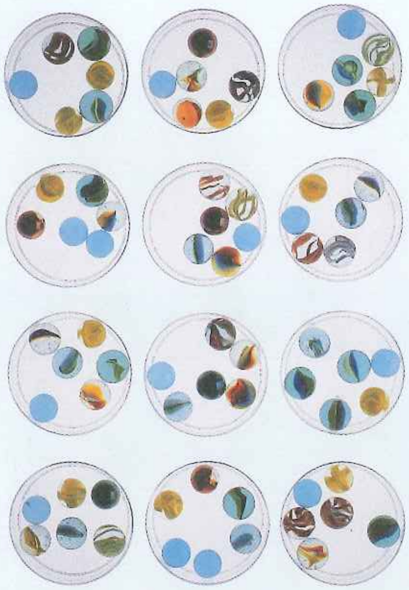


Figure 10.3 Counting Marbles
A dozen cups of marbles contain more than a dozen marbles. Similarly, a mole of molecules contains more than a mole of atoms.
Calculate ***How many atoms are in one mole of molecules if each molecule consists of six atoms?***

Suppose you want to determine how many atoms are in a mole of a compound. To do this, you must know how many atoms are in a representative particle of the compound. This number is determined from the chemical formula. Figure 10.3 illustrates this idea with marbles (atoms) in cups (molecules). The number of marbles in a dozen cups is (6×12), or 72 marbles. In the formula for carbon dioxide (CO_2), the subscripts show that one molecule of carbon dioxide is composed of three atoms: one carbon atom and two oxygen atoms. A mole of carbon dioxide contains Avogadro's number of CO_2 molecules. Each molecule contains three atoms, so a mole of carbon dioxide contains three times Avogadro's number of atoms. A molecule of carbon monoxide (CO) consists of two atoms, so a mole of carbon monoxide contains two times Avogadro's number of atoms.

To find the number of atoms in a given number of moles of a compound, you must first determine the number of representative particles. To convert the number of moles of a compound to the number of representative particles (molecules or formula units), multiply the number of moles by 6.02×10^{23} representative particles/1 mol. Then, multiply the number of representative particles by the number of atoms in each molecule or formula unit.

The Size of a Mole Perhaps you are wondering just how large a mole is. The SI unit, the mole, is not related to the small burrowing animal of the same name, shown in Figure 10.4. However, this little animal can help you appreciate the size of the number 6.02×10^{23}. Assume that an average animal-mole is 15 cm long, 5 cm tall, and has a mass of 145 g. Based on this information, the mass of 6.02×10^{23} animal-moles is 8.73×10^{22} kg. That means that the mass of Avogadro's number of animal-moles is equal to more than 60 times the combined mass of Earth's oceans. If spread over the entire surface of Earth, Avogadro's number of animal-moles would form a layer more than 8 million animal-moles thick. What about the length of 6.02×10^{23} animal-moles? If lined up end-to-end, 6.02×10^{23} animal-moles would stretch from Earth to the nearest star, Alpha Centauri, more than two million times. Are you beginning to understand how enormous Avogadro's number is?

Figure 10.4 A Mole of Moles
An average animal-mole has a mass of 145 g. The mass of 6.02×10^{23} animal-moles is 8.73×10^{22} kg.

CHEM TUTOR

Sample Problem 10.3

Converting Moles to Number of Atoms

Propane is a gas used for cooking and heating. How many atoms are in 2.12 mol of propane (C_3H_8)?

1 Analyze **List the knowns and the unknown.**
The desired conversion is moles ⟶ molecules ⟶ atoms.

KNOWNS
number of moles = 2.12 mol C_3H_8
1 mol C_3H_8 = 6.02×10^{23} molecules C_3H_8
1 molecule C_3H_8 = 11 atoms
(3 carbon atoms and 8 hydrogen atoms)

UNKNOWN
number of atoms = ? atoms

2 Calculate **Solve for the unknown.**

First, write the conversion factor to convert from moles to molecules.

$$\frac{6.02 \times 10^{23} \text{ molecules } C_3H_8}{1 \text{ mol } C_3H_8}$$

Write the conversion factor to convert from molecules to atoms.

$$\frac{11 \text{ atoms}}{1 \text{ molecule } C_3H_8}$$

Remember to write the conversion factors so that the unit in the denominator cancels the unit in the numerator of the previous factor.

Multiply the moles of C_3H_8 by the conversion factors.

$$2.12 \cancel{\text{mol } C_3H_8} \times \frac{6.02 \times 10^{23} \cancel{\text{molecules } C_3H_8}}{1 \cancel{\text{mol } C_3H_8}} \times \frac{11 \text{ atoms}}{1 \cancel{\text{molecule } C_3H_8}}$$

$$= \boxed{1.40 \times 10^{25} \text{ atoms}}$$

3 Evaluate **Does the result make sense?** There are 11 atoms in each molecule of propane and more than 2 mol of propane, so the answer should be more than 20 times Avogadro's number of propane molecules. The answer has three significant figures based on the three significant figures in the given measurement.

There are 3 atoms of carbon and 8 atoms of hydrogen in 1 molecule of propane.

5. How many atoms are in 1.14 mol of sulfur trioxide (SO_3)?

6. How many carbon atoms are in 2.12 mol of propane? How many hydrogen atoms are in 2.12 mol of propane?

Interpret Data

Carbon Atoms		Hydrogen Atoms		Mass Ratio
Number	Mass (amu)	Number	Mass (amu)	$\frac{\text{Mass carbon}}{\text{Mass hydrogen}}$
	12		1	$\frac{12\text{ amu}}{1\text{ amu}} = \frac{12}{1}$
	24 (2×12)		2 (2×1)	$\frac{24\text{ amu}}{2\text{ amu}} = \frac{12}{1}$
	120 (10×12)		10 (10×1)	$\frac{120\text{ amu}}{10\text{ amu}} = \frac{12}{1}$
	600 (50×12)		50 (50×1)	$\frac{600\text{ amu}}{50\text{ amu}} = \frac{12}{1}$
Avogadro's number	$(6.02 \times 10^{23}) \times (12)$	Avogadro's number	$(6.02 \times 10^{23}) \times (1)$	$\frac{(6.02 \times 10^{23}) \times (12)}{(6.02 \times 10^{23}) \times (1)} = \frac{12}{1}$

Table 10.2 An average carbon atom is 12 times heavier than an average hydrogen atom.

a. Read Tables What is the mass of 50 carbon atoms? What is the mass of 50 hydrogen atoms?

b. Apply Concepts What is the ratio of the mass of 500 carbon atoms to the mass of 500 hydrogen atoms?

c. Infer Do 36.0 kg of carbon atoms and 3.0 kg of hydrogen atoms contain the same number of atoms? Explain.

Hint: To answer part c, determine the mass ratio of carbon to hydrogen.

Molar Mass

How do you determine the molar mass of an element and of a compound?

Remember that the atomic mass of an element (the mass of a single atom) is expressed in atomic mass units (amu). The atomic masses are relative values based on the mass of the most common isotope of carbon (carbon-12). Table 10.2 shows that an average carbon atom (C) with an atomic mass of 12.0 amu is 12 times heavier than an average hydrogen atom (H) with an atomic mass of 1.0 amu. Therefore, 100 carbon atoms are 12 times heavier than 100 hydrogen atoms. In fact, any number of carbon atoms is 12 times heavier than the same number of hydrogen atoms. So 12.0 g of carbon atoms and 1.0 g of hydrogen atoms must contain the same number of atoms.

If you look at the atomic masses of the elements in the periodic table, you will notice that they are not whole numbers. For example, the atomic mass of carbon is not exactly 12 times the mass of hydrogen. Recall from Chapter 4 that this is because atomic masses are weighted average masses of the isotopes of each element.

The Mass of a Mole of an Element Quantities measured in grams are convenient for working in the laboratory, so chemists have converted the relative scale of masses of the elements in amu to a relative scale of masses in grams. **The atomic mass of an element expressed in grams is the mass of a mole of the element.** The mass of a mole of an element is its **molar mass.** For carbon, the molar mass is 12.0 g. For atomic hydrogen, the molar mass is 1.0 g. Figure 10.5 shows one mole of carbon, sulfur, and iron. Compare the molar masses in the figure to the atomic masses in your periodic table. Notice that the molar masses are rounded off to one place after the decimal point. All the examples and problems in this text use molar masses that are rounded off in this way. If your teacher uses a different rounding rule for molar masses, your answers to problems may differ slightly from the answers given in the text.

If you were to compare 12.0 g of carbon atoms with 16.0 g of oxygen atoms, you would find they contain the same number of atoms. The molar masses of any two elements must contain the same number of atoms. How many atoms are contained in the molar mass of an element? You already know. The molar mass of any element contains 1 mol or 6.02×10^{23} atoms of that element.

The mole can now be further defined as the amount of substance that contains the same number of representative particles as the number of atoms in 12.0 g of carbon-12. You know that 12.0 g is the molar mass of carbon-12, so 12.0 g of carbon is 1 mol of carbon atoms. The same relationship applies to hydrogen: 1.0 g of hydrogen is 1 mol of hydrogen atoms. Similarly, 24.3 g is the molar mass of magnesium, so 1 mol of magnesium (or 6.02×10^{23} atoms of magnesium) has a mass of 24.3 g. Molar mass is the mass of 1 mol of atoms of any element.

READING SUPPORT

Build Comprehension: ***Analogies*** You can buy small, medium, and large eggs. The size of the eggs doesn't affect how many eggs are in one dozen. Similarly, the size of the representative particles doesn't affect how many are in one mole. ***Can you think of another analogy to show the relationship between moles and the size of representative particles?***

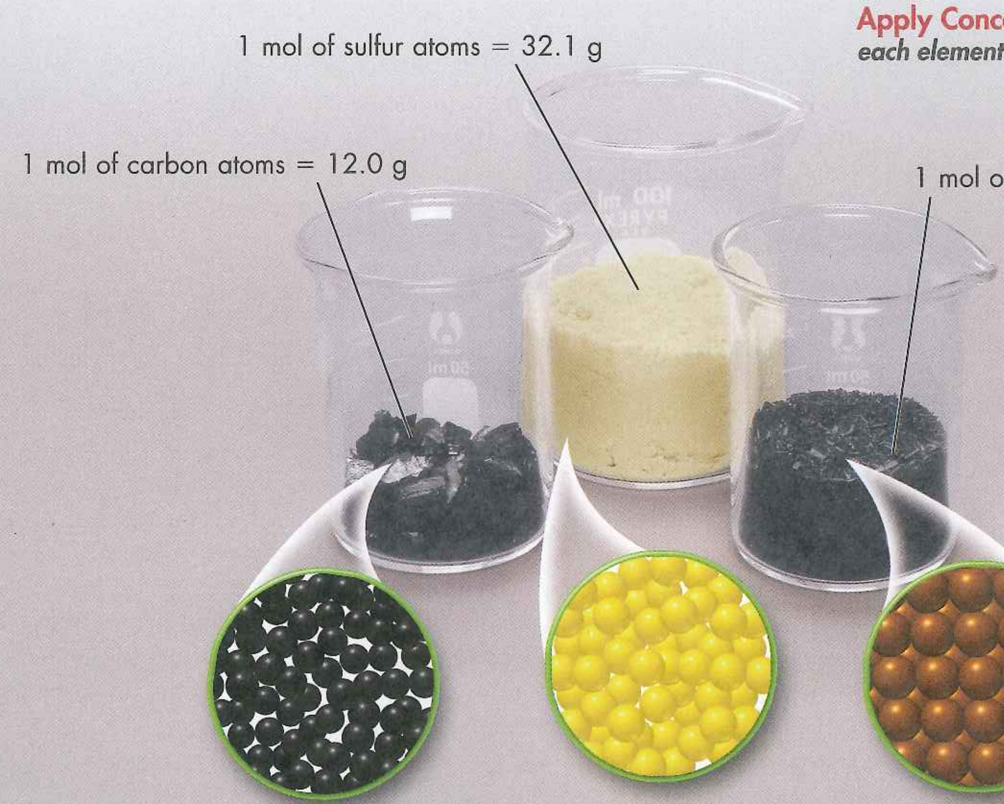


Figure 10.5 Molar Mass of an Element One mole of carbon, sulfur, and iron are shown.
Apply Concepts ***How many atoms of each element are present in each beaker?.***

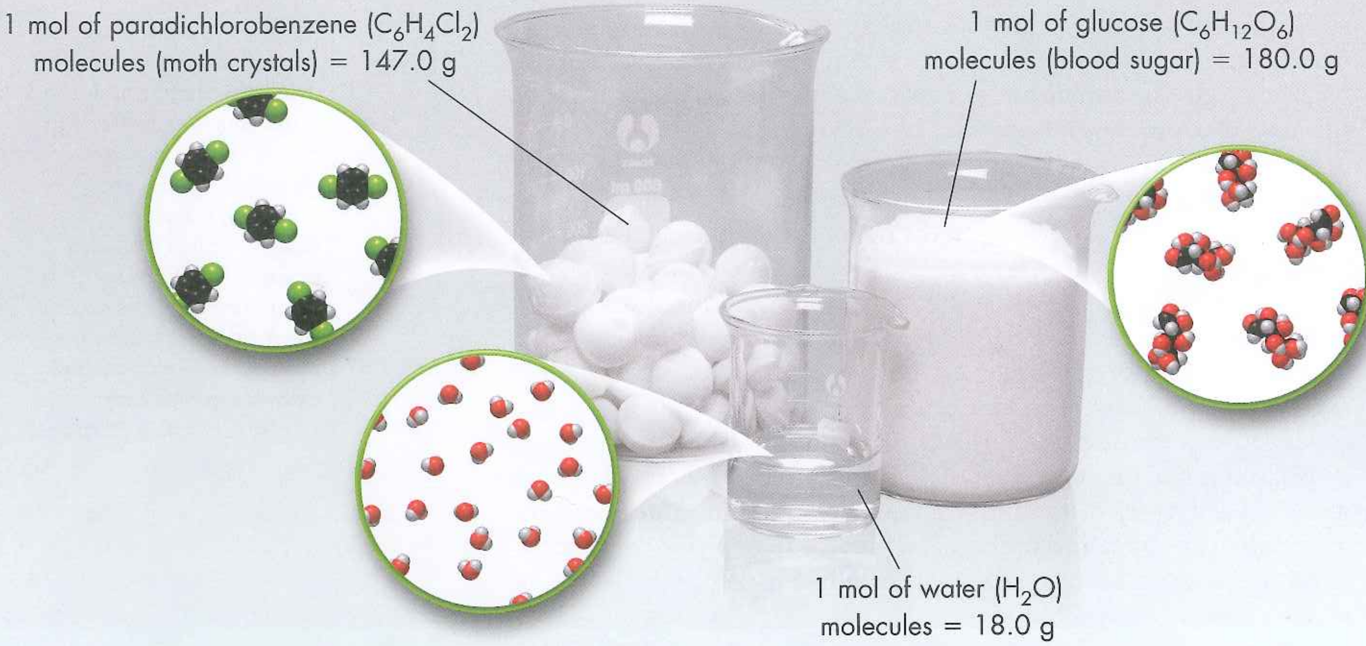


Figure 10.6 Molar Mass of a Compound
One mole is shown for each of three molecular compounds.
Infer ***How do you know that each sample contains Avogadro's number of molecules?***

KINETIC ART *See* **the molar masses of compounds** *animated online.*

The Mass of a Mole of a Compound To find the mass of a mole of a compound, you must know the formula of the compound. The formula of sulfur trioxide is SO_3. A molecule of SO_3 is composed of one atom of sulfur and three atoms of oxygen.

1 SO_3 molecule = 1 S atom + 3 O atoms

You can calculate the mass of a molecule of SO_3 by adding the atomic masses of the atoms making up the molecule. From the periodic table, the atomic mass of sulfur (S) is 32.1 amu. The mass of three atoms of oxygen is three times the atomic mass of a single oxygen atom (O): 3×16.0 amu = 48.0 amu. So, the molecular mass of SO_3 is 32.1 amu + 48.0 amu = 80.1 amu.

1 S atom + 3 O atoms = 1 SO_3 molecule

32.1 amu + 16.0 amu + 16.0 amu + 16.0 amu = 80.1 amu

Now substitute the unit grams for atomic mass units to find the molar mass of SO_3. The molar mass (g/mol) of any compound is the mass in grams of 1 mol of that compound. Thus, 1 mol of SO_3 has a mass of 80.1 g. This is the mass of 6.02×10^{23} molecules of SO_3.

To calculate the molar mass of a compound, find the number of grams of each element in one mole of the compound. Then add the masses of the elements in the compound. This method for calculating molar mass applies to any compound, molecular or ionic. The molar masses of paradichlorobenzene ($C_6H_4Cl_2$, 147.0 g), water (H_2O, 18.0 g), and glucose ($C_6H_{12}O_6$, 180.0 g) in Figure 10.6 were obtained in this way.

Sample Problem 10.4

Finding the Molar Mass of a Compound

The decomposition of hydrogen peroxide (H_2O_2) provides sufficient energy to launch a rocket. What is the molar mass of hydrogen peroxide?

KNOWNS

molecular formula = H_2O_2

mass of 1 mol H = 1.0 g H

mass of 1 mol O = 16.0 g O

UNKNOWN

molar mass = ? g/mol

1 Analyze **List the knowns and the unknown.** Convert moles of atoms to grams by using conversion factors (g/mol) based on the molar mass of each element. The sum of the masses of the elements is the molar mass.

2 Calculate **Solve for the unknown.**

Convert moles of hydrogen and oxygen to grams of hydrogen and oxygen.

$$2\ \cancel{\text{mol H}} \times \frac{1.0\ \text{g H}}{1\ \cancel{\text{mol H}}} = 2.0\ \text{g H}$$

$$2\ \cancel{\text{mol O}} \times \frac{16.0\ \text{g O}}{1\ \cancel{\text{mol O}}} = 32.0\ \text{g O}$$

One mole of H_2O_2 has 2 mol of H atoms and 2 mol of O atoms, so multiply the molar mass of each element by 2.

Add the results.

mass of 1 mol H_2O_2 = 2.0 g H + 32.0 g O = 34.0 g

molar mass of H_2O_2 = 34.0 g/mol

3 Evaluate **Does the result make sense?** The answer is the sum of two times the molar mass of hydrogen and oxygen (17.0 g/mol). The answer is expressed to the tenths place because the numbers being added are expressed to the tenths place.

One mole of PCl_3 has 1 mol of P atoms and 3 mol of Cl atoms.

7. Find the molar mass of PCl_3.

8. What is the mass of 1.00 mol of sodium hydrogen carbonate?

10.1 LessonCheck

9. **Review** What do you need to know to convert among the count, mass, and volume of something?

10. **Describe** How do chemists count the number of representative particles in a substance?

11. **Explain** How do you determine the molar mass of an element? How do you determine the molar mass of a compound?

12. **Calculate** If a dozen apples has a mass of 2.0 kg and 0.20 bushel is 1 dozen apples, how many bushels of apples are in 1.0 kg of apples?

13. **Calculate** How many moles is 1.50×10^{23} molecules of NH_3?

14. **Calculate** How many atoms are in 1.75 mol of $CHCl_3$?

15. **Calculate** What is the molar mass of $CaSO_4$?

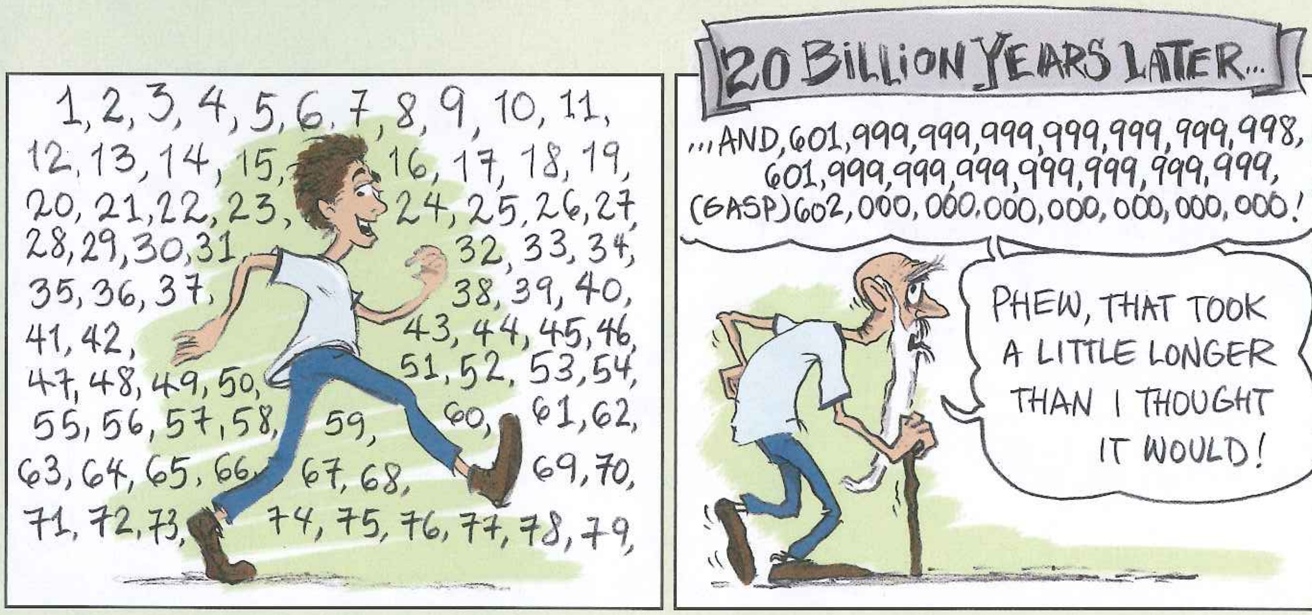
CHEMISTRY & YOU: EVERYDAY MATTER

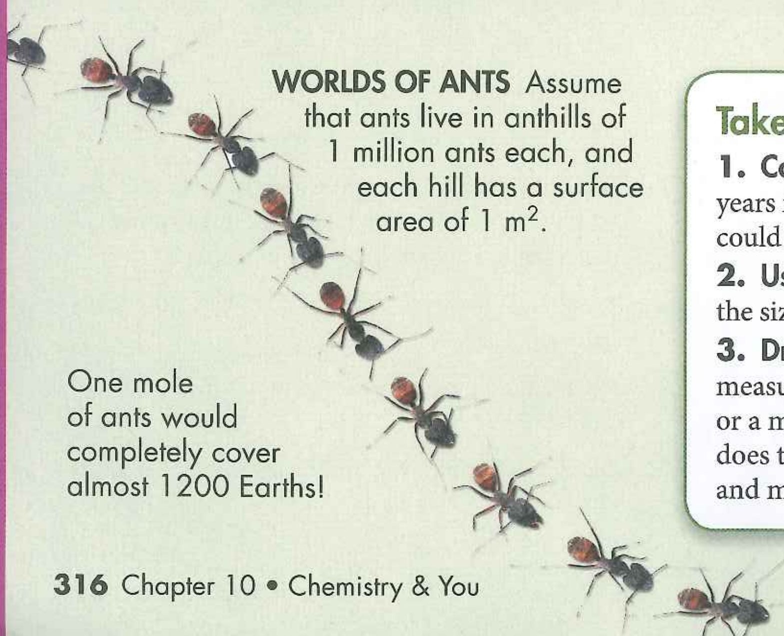
How Big Is a Mole?

The mole is an especially useful tool to chemists, because it allows them to express the number of representative particles of a substance in grams. For example, a 1 mol sample of carbon, which contains Avogadro's number of carbon atoms (6.02×10^{23}), has a mass of 12.0 g.

The mole is a huge quantity. Written out, Avogadro's number is 602,000,000,000,000,000,000,000. However, it may be difficult for you to comprehend exactly how big a mole is. Here are some interesting ways to visualize the size of a mole.

TOO BIG TO COUNT If you were able to count at the rate of 1 million numbers per second, it would take almost 20 billion years to count to 6.02×10^{23}.





WORLDS OF ANTS Assume that ants live in anthills of 1 million ants each, and each hill has a surface area of 1 m^2.

One mole of ants would completely cover almost 1200 Earths!

Take It Further

1. Calculate Show how to calculate the number of years it would take to count to Avogadro's number if you could count at the rate of 1 million numbers per second.

2. Use Models Develop your own concept to illustrate the size of Avogadro's number. Show your calculations.

3. Draw Conclusions At home, using a food scale, measure out a mole of table sugar (sucrose, $C_{12}H_{22}O_{11}$) or a mole of table salt (sodium chloride, NaCl). What does this measurement tell you about the size of atoms and molecules?

10.2 Mole–Mass and Mole–Volume Relationships

Key Questions

How do you convert the mass of a substance to the number of moles of the substance?

How do you convert the volume of a gas at STP to the number of moles of the gas?

Vocabulary

- Avogadro's hypothesis
- standard temperature and pressure (STP)
- molar volume

CHEMISTRY & YOU

Q: *How can you calculate the moles of a substance in a given mass or volume?* Guess how many pennies are in the container and win a prize! You decide to enter the contest, and you win. You estimated the thickness and diameter of a penny to find its approximate volume. Then you estimated the dimensions of the container to obtain its volume. You did the arithmetic and made your guess. In a similar way, chemists use the relationships between the mole and quantities such as mass, volume, and number of particles to solve problems in chemistry.

The Mole–Mass Relationship

How do you convert the mass of a substance to the number of moles of the substance?

In the previous lesson, you learned that the molar mass of any substance is the mass in grams of one mole of that substance. This definition applies to all substances—elements, molecular compounds, and ionic compounds. In some situations, however, the term *molar mass* may be unclear. For example, suppose you were asked for the molar mass of oxygen. How you answer this question depends on what you assume to be the representative particle. If you assume the oxygen in the question is molecular oxygen (O_2), then the molar mass is 32.0 g/mol (2×16.0 g/mol). If you assume that the question is asking for the mass of a mole of oxygen atoms (O), then the answer is 16.0 g/mol. You can avoid confusion such as this by using the formula of the substance, in this case, O_2 or O.

Suppose you need a given number of moles of a substance for a laboratory experiment. How can you measure this amount? Suppose instead that you obtain a certain mass of a substance in a laboratory experiment. How many moles is this? **Use the molar mass of an element or compound to convert between the mass of a substance and the moles of the substance.** The conversion factors for these calculations are based on the relationship: molar mass = 1 mol.

$$\frac{\text{molar mass}}{1\text{ mol}} \quad \text{and} \quad \frac{1\text{ mol}}{\text{molar mass}}$$

Sample Problem 10.5

Converting Moles to Mass

Items made out of aluminum, such as aircraft parts and cookware, are resistant to corrosion because the aluminum reacts with oxygen in the air to form a coating of aluminum oxide (Al_2O_3). This tough, resistant coating prevents any further corrosion. What is the mass, in grams, of 9.45 mol of aluminum oxide?

1 Analyze **List the known and the unknown.** The mass of the compound is calculated from the known number of moles of the compound. The desired conversion is moles ⟶ mass.

KNOWN

number of moles = 9.45 mol Al_2O_3

UNKNOWN

mass = ? g Al_2O_3

2 Calculate **Solve for the unknown.**

First, determine the mass of 1 mol of Al_2O_3.

$$2\ \cancel{\text{mol Al}} \times \frac{27.0\text{ g Al}}{1\ \cancel{\text{mol Al}}} = 54.0\text{ g Al}$$

$$3\ \cancel{\text{mol O}} \times \frac{16.0\text{ g O}}{1\ \cancel{\text{mol O}}} = 48.0\text{ g O}$$

$$1\text{ mol }Al_2O_3 = 54.0\text{ g Al} + 48.0\text{ g O} = 102.0\text{ g }Al_2O_3$$

Identify the conversion factor relating moles of Al_2O_3 to grams of Al_2O_3.

$$\frac{102.0\text{ g }Al_2O_3}{1\text{ mol }Al_2O_3}$$

Use the relationship: 1 mol Al_2O_3 = 102.0 g Al_2O_3.

Multiply the given number of moles by the conversion factor.

$$9.45\ \cancel{\text{mol }Al_2O_3} \times \frac{102.0\text{ g }Al_2O_3}{1\ \cancel{\text{mol }Al_2O_3}}$$

$$= 964\text{ g }Al_2O_3$$

3 Evaluate **Does the result make sense?** The number of moles of Al_2O_3 is approximately 10, and each has a mass of approximately 100 g. The answer should be close to 1000 g. The answer has been rounded to the correct number of significant figures.

16. Find the mass, in grams, of 4.52×10^{-3} mol $C_{20}H_{42}$.

17. Calculate the mass, in grams, of 2.50 mol of iron(II) hydroxide.

Start by determining the molar mass of each compound.

Sample Problem 10.6

Converting Mass to Moles

When iron is exposed to air, it corrodes to form red-brown rust. Rust is iron(III) oxide (Fe_2O_3). How many moles of iron(III) oxide are contained in 92.2 g of pure Fe_2O_3?

1 Analyze **List the known and the unknown.** The number of moles of the compound is calculated from the known mass of the compound. The conversion is mass $\longrightarrow$ moles.

KNOWN

mass $= 92.2\ g\ Fe_2O_3$

UNKNOWN

number of moles $= ?\ mol\ Fe_2O_3$

2 Calculate **Solve for the unknown.**

First, determine the mass of 1 mol of Fe_2O_3.

$$2\ \cancel{mol\ Fe} \times \frac{55.8\ g\ Fe}{1\ \cancel{mol\ Fe}} = 111.6\ g\ Fe$$

$$3\ \cancel{mol\ O} \times \frac{16.0\ g\ O}{1\ \cancel{mol\ O}} = 48.0\ g\ O$$

$$1\ mol\ Fe_2O_3 = 111.6\ g\ Fe + 48.0\ g\ O = 159.6\ g\ Fe_2O_3$$

Identify the conversion factor relating grams of Fe_2O_3 to moles of Fe_2O_3.

$$\frac{1\ mol\ Fe_2O_3}{159.6\ g\ Fe_2O_3}$$

Note that the known unit (g) is in the denominator and the unknown unit (mol) is in the numerator.

Multiply the given mass by the conversion factor.

$$92.2\ \cancel{g\ Fe_2O_3} \times \frac{1\ mol\ Fe_2O_3}{159.6\ \cancel{g\ Fe_2O_3}}$$

$$= 0.578\ mol\ Fe_2O_3$$

3 Evaluate **Does the result make sense?** The given mass (about 90 g) is slightly larger than the mass of one-half mole of Fe_2O_3 (about 80 g), so the answer should be slightly larger than one-half (0.5) mol.

18. Find the number of moles in 3.70×10^{-1} g of boron.

19. Calculate the number of moles in 75.0 g of dinitrogen trioxide.

Again, start by determining the molar mass of each substance.

The Mole–Volume Relationship

How do you convert the volume of a gas at STP to the number of moles of the gas?

CHEMISTRY & YOU

Q: *How can you calculate the moles of a substance in a given mass? How can you calculate the moles of a gas in a given volume at STP?*

Look back at Figure 10.6. Notice that the volumes of one mole of different solid and liquid substances are not the same. For example, the volumes of one mole of glucose (blood sugar) and one mole of paradichlorobenzene (moth crystals) are much larger than the volume of one mole of liquid water. What about the volumes of gases? Unlike liquids and solids, the volumes of moles of gases, measured under the same physical conditions, are much more predictable. Why is this?

Avogadro's Hypothesis In 1811, Amedeo Avogadro proposed a groundbreaking explanation. **Avogadro's hypothesis** states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. The particles that make up different gases are not the same size. However, the particles in all gases are so far apart that a collection of relatively large particles does not require much more space than the same number of relatively small particles. Whether the particles are large or small, large expanses of space exist between individual particles of gas, as shown in Figure 10.7.

The volume of a gas varies with a change in temperature or a change in pressure. Due to these variations with temperature and pressure, the volume of a gas is usually measured at a standard temperature and pressure. **Standard temperature and pressure (STP)** means a temperature of 0°C and a pressure of 101.3 kPa, or 1 atmosphere (atm). At STP, 1 mol, or 6.02×10^{23} representative particles, of any gas occupies a volume of 22.4 L. The quantity, 22.4 L, is called the **molar volume** of a gas.

Calculating the Volume and Moles of a Gas at STP The molar volume of a gas at STP is a useful quantity to chemists. **The molar volume is used to convert between the number of moles of gas and the volume of the gas at STP.** The conversion factors for these calculations are based on the relationship 22.4 L = 1 mol at STP.

$$\frac{22.4\ \text{L}}{1\ \text{mol}} \quad \text{and} \quad \frac{1\ \text{mol}}{22.4\ \text{L}}$$

You can use these conversion factors to convert a known number of moles of gas to the volume of the gas at STP. Similarly, you can convert a known volume of gas at STP to the number of moles of the gas.

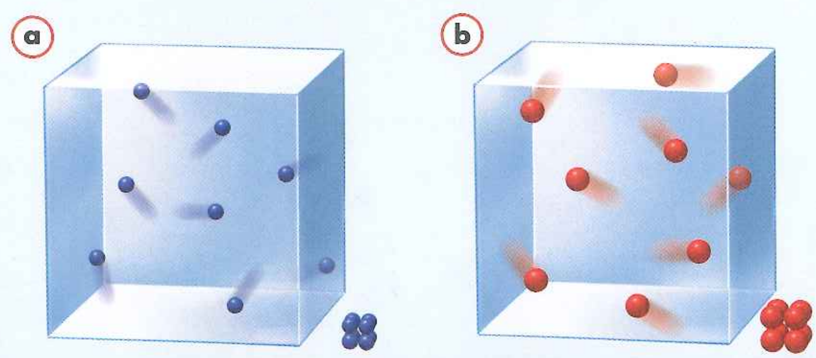


Figure 10.7 Volumes of Gases
In each container, the volume occupied by the gas molecules is small compared with the container's volume. **a.** The molecules in this container are small. **b.** This container can accommodate the same number of larger molecules.
Infer ***If the containers contained liquid molecules, and the molecules in container a were smaller than the molecules in container b, would both containers be able to accommodate the same number of molecules? Explain.***

SampleProblem 10.7

Calculating Gas Quantities at STP

Sulfur dioxide (SO_2) is a gas produced by burning coal. It is an air pollutant and one of the causes of acid rain. Determine the volume, in liters, of 0.60 mol SO_2 gas at STP.

1 Analyze List the knowns and the unknown. Since SO_2 is a gas, the volume at STP can be calculated from the known number of moles.

KNOWNS

number of moles = 0.60 mol SO_2

1 mol SO_2 = 22.4 L SO_2 at STP

UNKNOWN

volume = ? L SO_2

2 Calculate Solve for the unknown.

First, identify the conversion factor relating moles of SO_2 to volume of SO_2 at STP.

$$\frac{22.4\ \text{L}\ SO_2}{1\ \text{mol}\ SO_2}$$

The following relationship applies for gases at STP: 22.4 L = 1 mol.

Multiply the given number of moles by the conversion factor.

$$0.60\ \cancel{\text{mol}\ SO_2} \times \frac{22.4\ \text{L}\ SO_2}{1\ \cancel{\text{mol}\ SO_2}}$$

$$= \boxed{13\ \text{L}\ SO_2}$$

3 Evaluate Does the result make sense? One mole of any gas at STP has a volume of 22.4 L, so 0.60 mol should have a volume slightly larger than one half of a mole or 11.2 L. The answer should have two significant figures.

20. What is the volume of these gases at STP?

a. 3.20×10^{-3} mol CO_2

b. 3.70 mol N_2

c. 0.960 mol CH_4

21. At STP, how many moles are in these volumes of gases?

a. 67.2 L SO_2

b. 0.880 L He

c. 1.00×10^3 L C_2H_6

In Problem 20, convert from moles of gas to volume.

In Problem 21, convert from volume of gas to moles.

Calculating Molar Mass and Density A gas-filled balloon will either sink or float in the air depending on whether the density of the gas inside the balloon is greater or less than the density of the surrounding air. Different gases have different densities. Usually the density of a gas is measured in grams per liter (g/L) and at a specific temperature. The density of a gas at STP and the molar volume at STP (22.4 L/mol) can be used to calculate the molar mass of the gas. Similarly, the molar mass of a gas and the molar volume at STP can be used to calculate the density of a gas at STP.

You have now examined a mole in terms of particles, mass, and volume of gases at STP. Figure 10.8 summarizes these relationships and illustrates the importance of the mole.

Sample Problem 10.8

Calculating the Molar Mass of a Gas at STP

The density of a gaseous compound containing carbon and oxygen is found to be 1.964 g/L at STP. What is the molar mass of the compound?

KNOWNS

density = 1.964 g/L

1 mol of gas at STP = 22.4 L

UNKNOWN

molar mass = ? g/mol

1 Analyze **List the knowns and the unknown.** The molar mass of the compound is calculated from the known density of the compound and the molar volume at STP.

2 Calculate **Solve for the unknown.**

First, identify the conversion factor needed to convert density to molar mass.

$$\frac{22.4\text{ L}}{1\text{ mol}}$$

Use the density and molar volume at STP to calculate the molar mass.

$$\text{molar mass} = \frac{\text{g}}{\text{mol}} = \frac{\text{g}}{\text{L}} \times \frac{22.4\text{ L}}{1\text{ mol}}$$

Multiply the given density by the conversion factor.

$$\frac{1.964\text{ g}}{1\ \cancel{\text{L}}} \times \frac{22.4\ \cancel{\text{L}}}{1\text{ mol}}$$

$$= 44.0\text{ g/mol}$$

3 Evaluate **Does the result make sense?** The ratio of the calculated mass (44.0 g) to the volume (22.4 L) is about 2, which is close to the known density. The answer should have three significant figures.

22. A gaseous compound composed of sulfur and oxygen has a density of 3.58 g/L at STP. What is the molar mass of this gas?

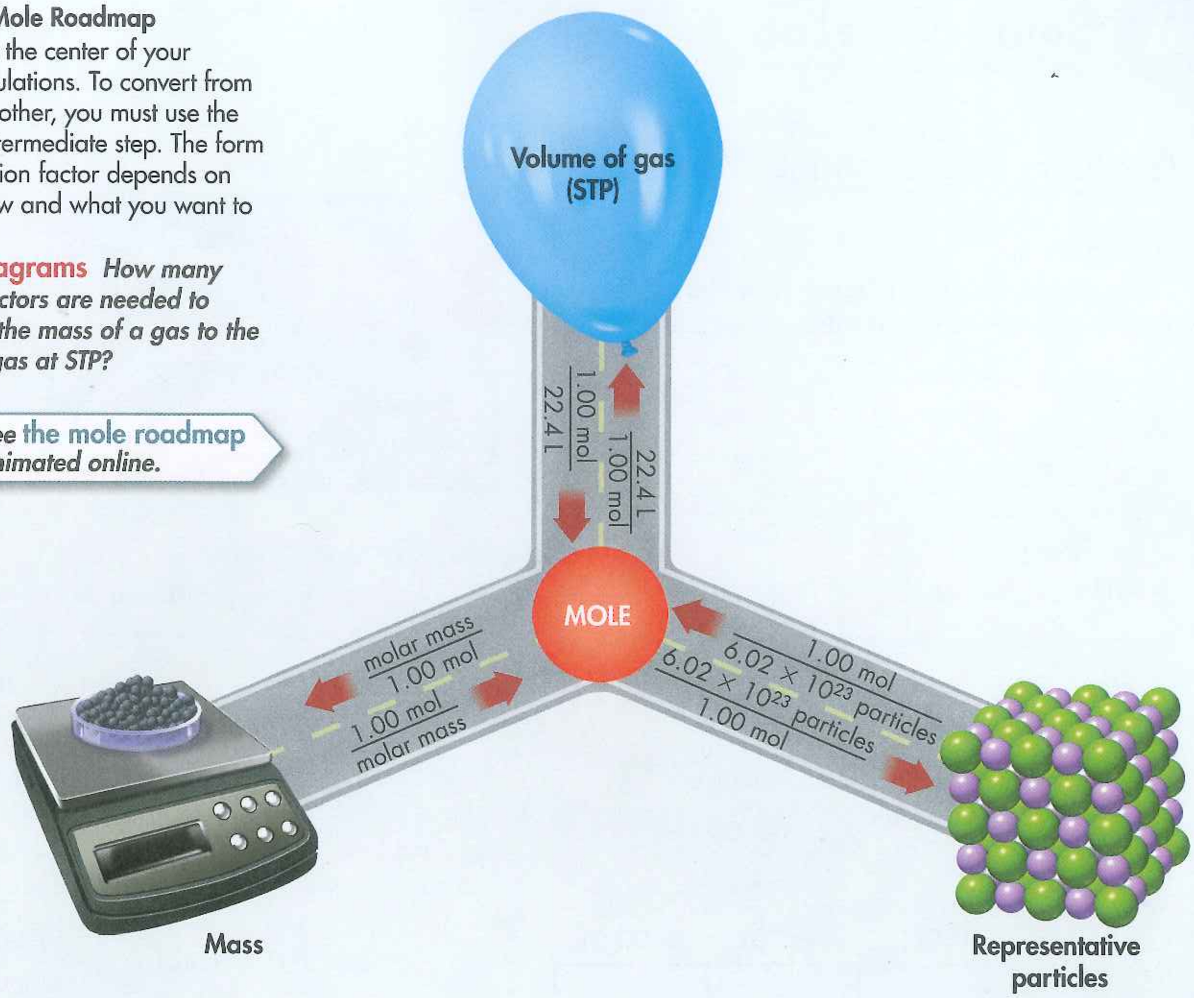
In Problem 22, use the density and molar volume of the gas at STP to calculate the molar mass.

23. What is the density of krypton gas at STP?

To do Problem 23, first find the molar mass of krypton. Use the molar mass and the molar volume at STP to calculate the density at STP.

Figure 10.8 Mole Roadmap
The mole is at the center of your chemical calculations. To convert from one unit to another, you must use the mole as an intermediate step. The form of the conversion factor depends on what you know and what you want to calculate.
Interpret Diagrams *How many conversion factors are needed to convert from the mass of a gas to the volume of a gas at STP?*

See **the mole roadmap** *animated online.*



10.2 LessonCheck

24. Describe How do you convert between the mass and the number of moles of a substance?

25. Describe How do you convert between the volume of a gas at STP and the number of moles of the gas?

26. Calculate How many grams are in 5.66 mol of $CaCO_3$?

27. Calculate Find the number of moles in 508 g of ethanol (C_2H_6O).

28. Calculate What is the volume, in liters, of 1.50 mol Cl_2 at STP?

29. Apply Concepts Three balloons filled with three different gaseous compounds each have a volume of 22.4 L at STP. Do these balloons have the same mass or contain the same number of molecules? Explain.

30. Calculate The density of an elemental gas is 1.7824 g/L at STP. What is the molar mass of the element?

31. Analyze Data The densities of gases A, B, and C at STP are 1.25 g/L, 2.86 g/L, and 0.714 g/L, respectively. Calculate the molar mass of each substance. Identify each substance as ammonia (NH_3), sulfur dioxide (SO_2), chlorine (Cl_2), nitrogen (N_2), or methane (CH_4).

BIGIDEA
THE MOLE AND QUANTIFYING MATTER

32. A chemist collects 2.94 L of carbon monoxide (CO) gas at STP during an experiment. Explain how she can determine the mass of gas that she collected. Why is the mole important for this calculation?

Small-Scale Lab

Counting by Measuring Mass

Purpose

To determine the mass of several samples of chemical compounds and use the data to count atoms

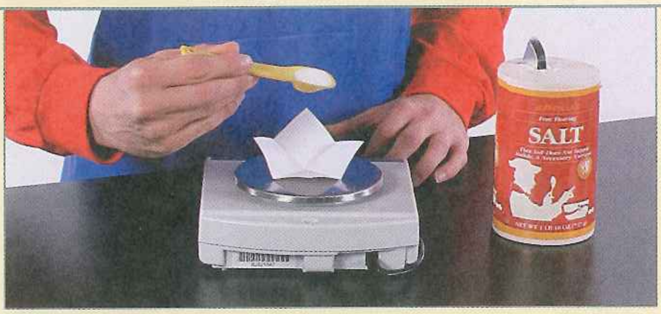
Materials

- H_2O, NaCl, and $CaCO_3$
- plastic spoon
- weighing paper
- balance

Procedure

Measure the mass of one level teaspoon of water (H_2O), sodium chloride (NaCl), and calcium carbonate ($CaCO_3$). Make a table similar to the one below.

	$H_2O(l)$	NaCl(s)	$CaCO_3(s)$
Mass (g)			
Molar mass (g/mol)			
Moles of each compound			
Moles of each element			
Atoms of each element			



Analyze and Conclude

1. Calculate Determine the number of moles of H_2O contained in one level teaspoon.

$$\text{moles of } H_2O = \text{g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0 \text{ g } H_2O}$$

Repeat for the remaining compounds. Use the periodic table to calculate the molar masses of NaCl and $CaCO_3$.

2. Calculate Determine the number of moles of each element present in the teaspoon-sized sample of H_2O.

$$\text{moles of H} = \text{mol } H_2O \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O}$$

Repeat for the other compounds in your table.

3. Calculate Determine the number of atoms of each element present in the teaspoon-sized sample of H_2O.

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

Repeat for the other compounds in your table.

4. Analyze Data Which of the three teaspoon-sized samples contains the greatest number of moles of molecules or formula units?

5. Analyze Data Which of the three compounds contains the greatest number of atoms?

You're the Chemist

1. Design an Experiment Can you count by measuring volume? Design and carry out an experiment to do it.

2. Design an Experiment Design an experiment that will determine the number of atoms of calcium, carbon, and oxygen it takes to write your name on the chalkboard with a piece of chalk. Assume chalk is 100 percent calcium carbonate, $CaCO_3$.

10.3 Percent Composition and Chemical Formulas

CHEMISTRY & YOU

Q: *What does the percent composition of a compound tell you?* A tag sewn into the seam of a shirt usually tells you what fibers were used to make the cloth and the percent of each. It helps to know the percents of the components in the shirt because they affect how warm the shirt is, whether it will need to be ironed, and how it should be cleaned. Similarly, in chemistry it is important to know the percents of the elements in a compound.

Key Questions

How do you calculate the percent composition of a compound?

How can you calculate the empirical formula of a compound?

How does the molecular formula of a compound compare with the empirical formula?

Vocabulary

- percent composition
- empirical formula

Percent Composition of a Compound

How do you calculate the percent composition of a compound?

In lawn care, the relative amount, or the percent, of each nutrient in fertilizer is important. In spring, you may use a fertilizer that has a high percent of nitrogen to "green" the grass. In fall, you may want to use a fertilizer with a higher percent of potassium to strengthen the root system. Knowing the relative amounts of the components of a mixture or compound is often useful.

The relative amounts of the elements in a compound are expressed as the **percent composition** or the percent by mass of each element in the compound. As shown in Figure 10.9, the percent composition of potassium chromate, K_2CrO_4, is K = 40.3%, Cr = 26.8%, and O = 32.9%. These percents must total 100% (40.3% + 26.8% + 32.9% = 100%). The percent composition of a compound is always the same.

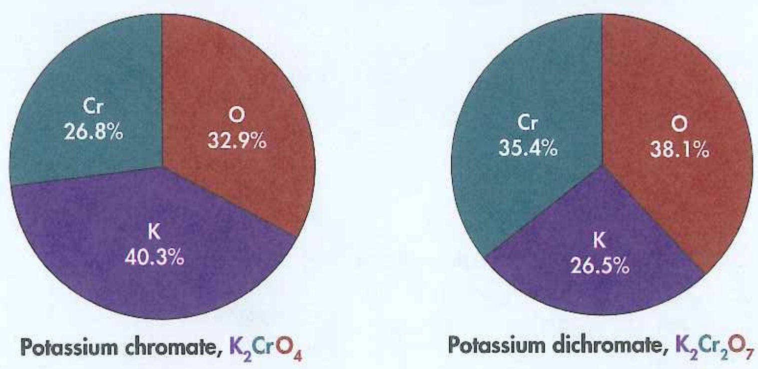


Figure 10.9 Percent Composition
Potassium chromate (K_2CrO_4) is composed of 40.3% potassium, 26.8% chromium, and 32.9% oxygen.
Compare ***How does this percent composition differ from the percent composition of potassium dichromate ($K_2Cr_2O_7$), a compound composed of the same three elements?***

Percent Composition From Mass Data If you know the relative masses of each element in a compound, you can calculate the percent composition of the compound. **The percent by mass of an element in a compound is the number of grams of the element divided by the mass in grams of the compound, multiplied by 100%.**

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

Sample Problem 10.9

Calculating Percent Composition From Mass Data

When a 13.60-g sample of a compound containing only magnesium and oxygen is decomposed, 5.40 g of oxygen is obtained. What is the percent composition of this compound?

1 Analyze List the knowns and the unknowns. The percent by mass of an element in a compound is the mass of that element divided by the mass of the compound multiplied by 100%.

KNOWNS

mass of compound = 13.60 g

mass of oxygen = 5.40 g O

mass of magnesium = 13.60 g − 5.40 g O = 8.20 g Mg

UNKNOWNS

percent by mass of Mg = ?% Mg

percent by mass of O = ?% O

2 Calculate Solve for the unknowns.

Determine the percent by mass of Mg in the compound.

$$\% \text{Mg} = \frac{\text{mass of Mg}}{\text{mass of compound}} \times 100\% = \frac{8.20\,\cancel{g}}{13.60\,\cancel{g}} \times 100\%$$

$$= 60.3\% \text{ Mg}$$

Determine the percent by mass of O in the compound.

$$\% \text{O} = \frac{\text{mass of O}}{\text{mass of compound}} \times 100\% = \frac{5.40\,\cancel{g}}{13.60\,\cancel{g}} \times 100\%$$

$$= 39.7\% \text{ O}$$

3 Evaluate Does the result make sense? The percents of the elements add up to 100%.

$$60.3\% + 39.7\% = 100\%$$

33. A compound is formed when 9.03 g Mg combines completely with 3.48 g N. What is the percent composition of this compound?

34. When a 14.2-g sample of mercury(II) oxide is decomposed into its elements by heating, 13.2 g Hg is obtained. What is the percent composition of the compound?

Percent Composition From the Chemical Formula You can also calculate the percent composition of a compound using its chemical formula. The subscripts in the formula are used to calculate the mass of each element in a mole of that compound. Using the individual masses of the elements and the molar mass, you can calculate the percent by mass of each element.

Learn more about percent composition *online.*

CONCEPTS IN ACTION

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

CHEM TUTOR

SampleProblem 10.10

Calculating Percent Composition From a Formula

Propane (C_3H_8), the fuel commonly used in gas grills, is one of the compounds obtained from petroleum. Calculate the percent composition of propane.

1 Analyze List the knowns and the unknowns. Calculate the percent by mass of each element by dividing the mass of that element in one mole of the compound by the molar mass of the compound and multiplying by 100%.

KNOWNS

mass of C in 1 mol C_3H_8 = 3 mol × 12.0 g/mol = 36.0 g
mass of H in 1 mol C_3H_8 = 8 mol × 1.0 g/mol = 8.0 g
molar mass of C_3H_8 = 36.0 g/mol + 8.0 g/mol = 44.0 g/mol

UNKNOWNS

percent by mass of C = ?% C
percent by mass of H = ?% H

2 Calculate Solve for the unknowns.

Determine the percent by mass of C in C_3H_8.

$$\%C = \frac{\text{mass of C in 1 mol } C_3H_8}{\text{molar mass of } C_3H_8} \times 100\% = \frac{36.0\cancel{g}}{44.0\cancel{g}} \times 100\%$$
$$= 81.8\%\ C$$

Determine the percent by mass of H in C_3H_8.

$$\%H = \frac{\text{mass of H in 1 mol } C_3H_8}{\text{molar mass of } C_3H_8} \times 100\% = \frac{8.0\cancel{g}}{44.0\cancel{g}} \times 100\%$$
$$= 18\%\ H$$

3 Evaluate Does the result make sense? The percents of the elements add up to 100% when the answers are expressed to two significant figures (82% + 18% = 100%).

35. Calculate the percent by mass of nitrogen in these fertilizers.
a. NH_3
b. NH_4NO_3

36. Calculate the percent composition of these compounds.
a. ethane (C_2H_6)
b. sodium hydrogen sulfate ($NaHSO_4$)

Quick Lab

Purpose To measure the percent of water in a series of crystalline compounds called hydrates

Materials

- 3 medium-sized test tubes
- balance
- spatula
- hydrated compounds of copper(II) sulfate, calcium chloride, and sodium sulfate
- test tube holder
- gas burner

Percent Composition

Procedure

1. Label each test tube with the name of a compound. Measure and record the masses.

2. Add 2–3 g of each compound (a good-sized spatula full) to the appropriately labeled test tube. Measure and record the mass of each test tube and the compound.

3. Using a test tube holder, hold one of the tubes at a 45° angle and gently heat its contents over the burner, slowly passing it in and out of the flame. Note any change in the appearance of the solid compound.

4. As moisture begins to condense in the upper part of the test tube, gently heat the entire length of the tube. Continue heating until all of the moisture is driven from the tube. This process may take 2–3 minutes. Repeat Steps 3 and 4 for the other two tubes.

5. Allow each tube to cool. Then measure and record the mass of each test tube and the heated compound.

Analyze and Conclude

1. Organize Data Set up a data table so that you can subtract the mass of the empty tube from the mass of the compound and the test tube, both before and after heating.

2. Calculate Find the difference between the mass of each compound before and after heating. This difference represents the amount of water lost by the hydrated compound due to heating.

3. Calculate Determine the percent by mass of water lost by each compound.

4. Analyze Data Which compound lost the greatest percent by mass of water? Which compound lost the smallest percent by mass of water?

CHEMISTRY & YOU

Q: *What information can you get from the percent composition of a compound?*

Percent Composition as a Conversion Factor You can use percent composition to calculate the number of grams of any element in a specific mass of a compound. To do this, multiply the mass of the compound by a conversion factor based on the percent composition of the element in the compound. In Sample Problem 10.10, you found that propane is 81.8 percent carbon and 18 percent hydrogen. That means that in a 100-g sample of propane, you would have 81.8 g of carbon and 18 g of hydrogen. You can use the following conversion factors to solve for the mass of carbon or hydrogen contained in a specific amount of propane.

$$\frac{81.8 \text{ g C}}{100 \text{ g } C_3H_8} \quad \text{and} \quad \frac{18 \text{ g H}}{100 \text{ g } C_3H_8}$$

Sample Problem 10.11

Calculating the Mass of an Element in a Compound Using Percent Composition

Calculate the mass of carbon and the mass of hydrogen in 82.0 g of propane (C_3H_8).

1 Analyze List the known and the unknowns. Use the conversion factors based on the percent composition of propane to make the following conversions: grams $C_3H_8 \longrightarrow$ grams C and grams $C_3H_8 \longrightarrow$ grams H.

KNOWN	UNKNOWNS
mass of C_3H_8 = 82.0 g	mass of carbon = ? g C
	mass of hydrogen = ? g H

2 Calculate Solve for the unknowns.

To calculate the mass of C, first write the conversion factor to convert from mass of C_3H_8 to mass of C.	$\dfrac{81.8\text{ g C}}{100\text{ g }C_3H_8}$	From Sample Problem 10.10, the percent by mass of C in C_3H_8 is 81.8%.
Multiply the mass of C_3H_8 by the conversion factor.	$82.0\ \cancel{\text{g }C_3H_8} \times \dfrac{81.8\text{ g C}}{100\ \cancel{\text{g }C_3H_8}} = \boxed{67.1\text{ g C}}$	
To calculate the mass of H, first write the conversion factor to convert from mass of C_3H_8 to mass of H.	$\dfrac{18\text{ g H}}{100\text{ g }C_3H_8}$	From Sample Problem 10.10, the percent by mass of H in C_3H_8 is 18%.
Multiply the mass of C_3H_8 by the conversion factor.	$82.0\ \cancel{\text{g }C_3H_8} \times \dfrac{18\text{ g H}}{100\ \cancel{\text{g }C_3H_8}} = \boxed{15\text{ g H}}$	

3 Evaluate Does the result make sense? The sum of the two masses equals 82 g, the sample size, to two significant figures (67 g C + 15 g H = 82 g C_3H_8).

37. Calculate the grams of nitrogen in 125 g of each fertilizer.

a. NH_3

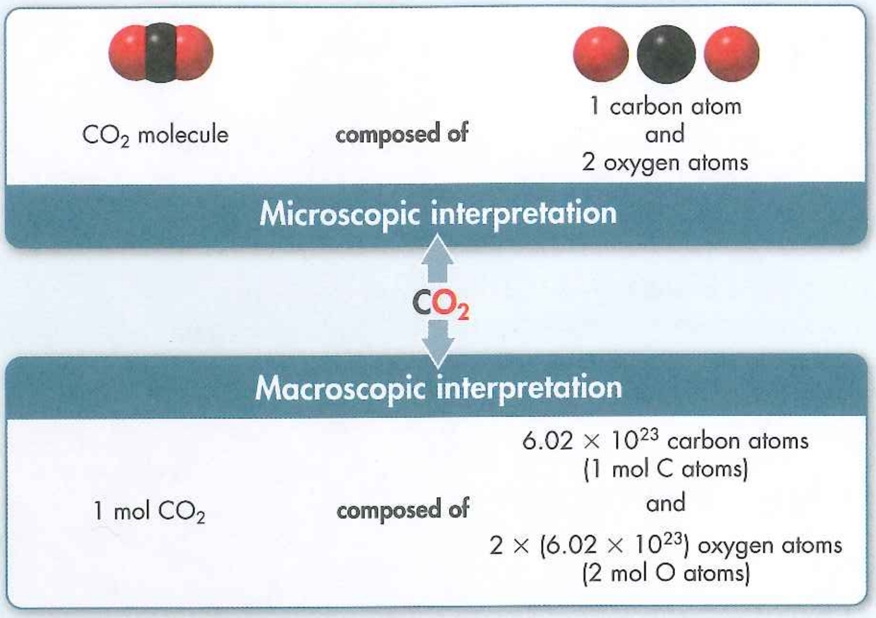
b. NH_4NO_3

38. Calculate the mass of hydrogen in each of the following compounds:

a. 350 g ethane (C_2H_6)

b. 20.2 g sodium hydrogen sulfate ($NaHSO_4$)

Figure 10.10 Interpreting Formulas A formula can be interpreted on a microscopic level in terms of atoms or on a macroscopic level in terms of moles of atoms.



Empirical Formulas

How can you calculate the empirical formula of a compound?

A useful formula for cooking rice is to use one cup of rice and two cups of water. If you needed twice the amount of cooked rice, you would need two cups of rice and four cups of water. The formulas for some compounds also show a basic ratio of elements. Multiplying that ratio by any factor can produce the formulas for other compounds.

The **empirical formula** of a compound gives the lowest whole-number ratio of the atoms or moles of the elements in a compound. Figure 10.10 shows that empirical formulas may be interpreted at the microscopic (atomic) or macroscopic (molar) level. An empirical formula may or may not be the same as a molecular formula. For example, the lowest ratio of hydrogen to oxygen in hydrogen peroxide is 1:1. Thus the empirical formula of hydrogen peroxide is HO. The molecular formula of hydrogen peroxide, H_2O_2, has twice the number of atoms as the empirical formula. Notice that the ratio of hydrogen to oxygen is still the same, 1:1. The molecular formula tells the actual number of each kind of atom present in a molecule of the compound. For carbon dioxide, the empirical and molecular formulas are the same—CO_2. Figure 10.11 shows two compounds of carbon and hydrogen having the same empirical formula (CH) but different molecular formulas.

The percent composition of a compound can be used to calculate the empirical formula of that compound. The percent composition tells the ratio of masses of the elements in a compound. The ratio of masses can be changed to a ratio of moles by using conversion factors based on the molar mass of each element. The mole ratio is then reduced to the lowest whole-number ratio to obtain the empirical formula of the compound.

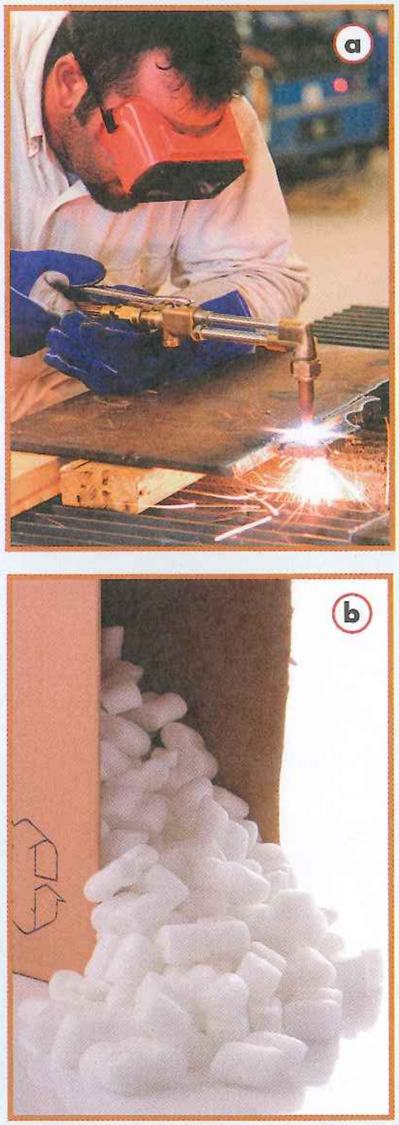


Figure 10.11 Compounds With the Same Empirical Formula
Two different compounds can have the same empirical formula. **a.** Ethyne (C_2H_2), also called acetylene, is a gas used in welders' torches. **b.** Styrene (C_8H_8) is used in making polystyrene.
Calculate ***What is the empirical formula of ethyne and styrene?***

Sample Problem 10.12

Determining the Empirical Formula of a Compound

A compound is analyzed and found to contain 25.9% nitrogen and 74.1% oxygen. What is the empirical formula of the compound?

1 Analyze **List the knowns and the unknown.** The percent composition gives the ratio of the mass of nitrogen atoms to the mass of oxygen atoms in the compound. Change the ratio of masses to a ratio of moles and reduce this ratio to the lowest whole-number ratio.

KNOWNS

percent by mass of N = 25.9% N

percent by mass of O = 74.1% O

UNKNOWN

empirical formula = $N_?O_?$

2 Calculate **Solve for the unknown.**

Convert the percent by mass of each element to moles.

$$25.9\,\cancel{g\,N} \times \frac{1\text{ mol N}}{14.0\,\cancel{g\,N}} = 1.85\text{ mol N}$$

$$74.1\,\cancel{g\,O} \times \frac{1\text{ mol O}}{16.0\,\cancel{g\,O}} = 4.63\text{ mol O}$$

The mole ratio of N to O is $N_{1.85}O_{4.63}$.

Percent means "parts per 100," so 100.0 g of the compound contains 25.9 g N and 74.1 g O.

Divide each molar quantity by the smaller number of moles to get 1 mol for the element with the smaller number of moles.

$$\frac{1.85\text{ mol N}}{1.85} = 1\text{ mol N}$$

$$\frac{4.63\text{ mol O}}{1.85} = 2.50\text{ mol O}$$

The mole ratio of N to O is now $N_1O_{2.5}$.

Multiply each part of the ratio by the smallest whole number that will convert both subscripts to whole numbers.

$$1\text{ mol N} \times 2 = 2\text{ mol N}$$
$$2.5\text{ mol O} \times 2 = 5\text{ mol O}$$

The empirical formula is N_2O_5.

3 Evaluate **Does the result make sense?** The subscripts are whole numbers, and the percent composition of this empirical formula equals the percents given in the original problem.

39. Calculate the empirical formula of each compound.

a. 94.1% O, 5.9% H

b. 67.6% Hg, 10.8% S, 21.6% O

40. 1,6-diaminohexane is used to make nylon. What is the empirical formula of this compound if its percent composition is 62.1% C, 13.8% H, and 24.1% N?

Interpret Data

Comparison of Empirical and Molecular Formulas		
Formula (name)	**Classification of formula**	**Molar mass (g/mol)**
CH	Empirical	13
C_2H_2 (ethyne)	Molecular	26 (2 × 13)
C_6H_6 (benzene)	Molecular	78 (6 × 13)
CH_2O (methanal)	Empirical and molecular	30
$C_2H_4O_2$ (ethanoic acid)	Molecular	60 (2 × 30)
$C_6H_{12}O_6$ (glucose)	Molecular	180 (6 × 30)

Table 10.3 Different compounds can have the same empirical formula.
a. Read Tables What is the molar mass of benzene, C_6H_6?
b. Interpret Tables Which compounds in the table have the empirical formula CH_2O?
c. Explain Why is the molar mass of glucose ($C_6H_{12}O_6$) equal to six times the molar mass of methanal (CH_2O)?

Hint: How is the formula $C_6H_{12}O_6$ related to the formula CH_2O?

Molecular Formulas

How does the molecular formula of a compound compare with the empirical formula?

Look at the compounds listed in Table 10.3. Ethyne and benzene have the same empirical formula—CH. Methanal, ethanoic acid, and glucose, shown in Figure 10.12, have the same empirical formula—CH_2O. Notice that the molar masses of the compounds in these two groups are simple whole-number multiples of the molar masses of the empirical formulas, CH and CH_2O. **The molecular formula of a compound is either the same as its experimentally determined empirical formula, or it is a simple whole-number multiple of its empirical formula.**

Once you have determined the empirical formula of a compound, you can determine its molecular formula, if you know the compound's molar mass. A chemist often uses an instrument called a mass spectrometer to determine molar mass. The compound is broken into charged fragments (ions) that travel through a magnetic field. The magnetic field deflects the particles from their straight-line paths. The mass of the compound is determined from the amount of deflection experienced by the particles.

You can calculate the empirical formula mass (efm) of a compound from its empirical formula. This is simply the molar mass of the empirical formula. Then you can divide the experimentally determined molar mass by the empirical formula mass. This quotient gives the number of empirical formula units in a molecule of the compound and is the multiplier to convert the empirical formula to the molecular formula.

Figure 10.12 Compounds With the Empirical Formula CH_2O
Methanal (formaldehyde), ethanoic acid (acetic acid), and glucose have the same empirical formula.
Apply Concepts ***How could you easily obtain the molar mass of ethanoic acid using the molar mass of methanal?***

Sample Problem 10.13

Finding the Molecular Formula of a Compound

Calculate the molecular formula of a compound whose molar mass is 60.0 g/mol and empirical formula is CH_4N.

KNOWNS
empirical formula = CH_4N
molar mass = 60.0 g/mol

UNKNOWN
molecular formula = $C_?H_?N_?$

1 Analyze **List the knowns and the unknown.** Divide the molar mass by the empirical formula mass to obtain a whole number. Multiply the empirical formula subscripts by this value to get the molecular formula.

2 Calculate **Solve for the unknown.**

First, calculate the empirical formula mass.

$$\text{efm of } CH_4N = 12.0\ \text{g/mol} + 4(1.0\ \text{g/mol}) + 14.0\ \text{g/mol} = 30.0\ \text{g/mol}$$

Divide the molar mass by the empirical formula mass.

$$\frac{\text{molar mass}}{\text{efm}} = \frac{60.0\ \cancel{\text{g/mol}}}{30.0\ \cancel{\text{g/mol}}} = 2$$

Multiply the formula subscripts by this value.

$$(CH_4N) \times 2 = C_2H_8N_2$$

3 Evaluate **Does the result make sense?** The molecular formula has the molar mass of the compound.

41. What is the molecular formula of a compound with the empirical formula CClN and a molar mass of 184.5 g/mol?

42. Find the molecular formula of ethylene glycol, which is used as antifreeze. The molar mass is 62.0 g/mol, and the empirical formula is CH_3O.

10.3 LessonCheck

43. Review How do you calculate the percent by mass of an element in a compound?

44. Identify What information can you use to calculate the empirical formula of a compound?

45. Explain How is the molecular formula of a compound related to its empirical formula?

46. Calculate Determine the percent composition of the compound that forms when 222.6 g N combines completely with 77.4 g O.

47. Calculate Find the percent composition of calcium acetate, $Ca(C_2H_3O_2)_2$.

48. Calculate Using the results of Problem 47, calculate the grams of hydrogen in 124 g of $Ca(C_2H_3O_2)_2$.

BIGIDEA
THE MOLE AND QUANTIFYING MATTER

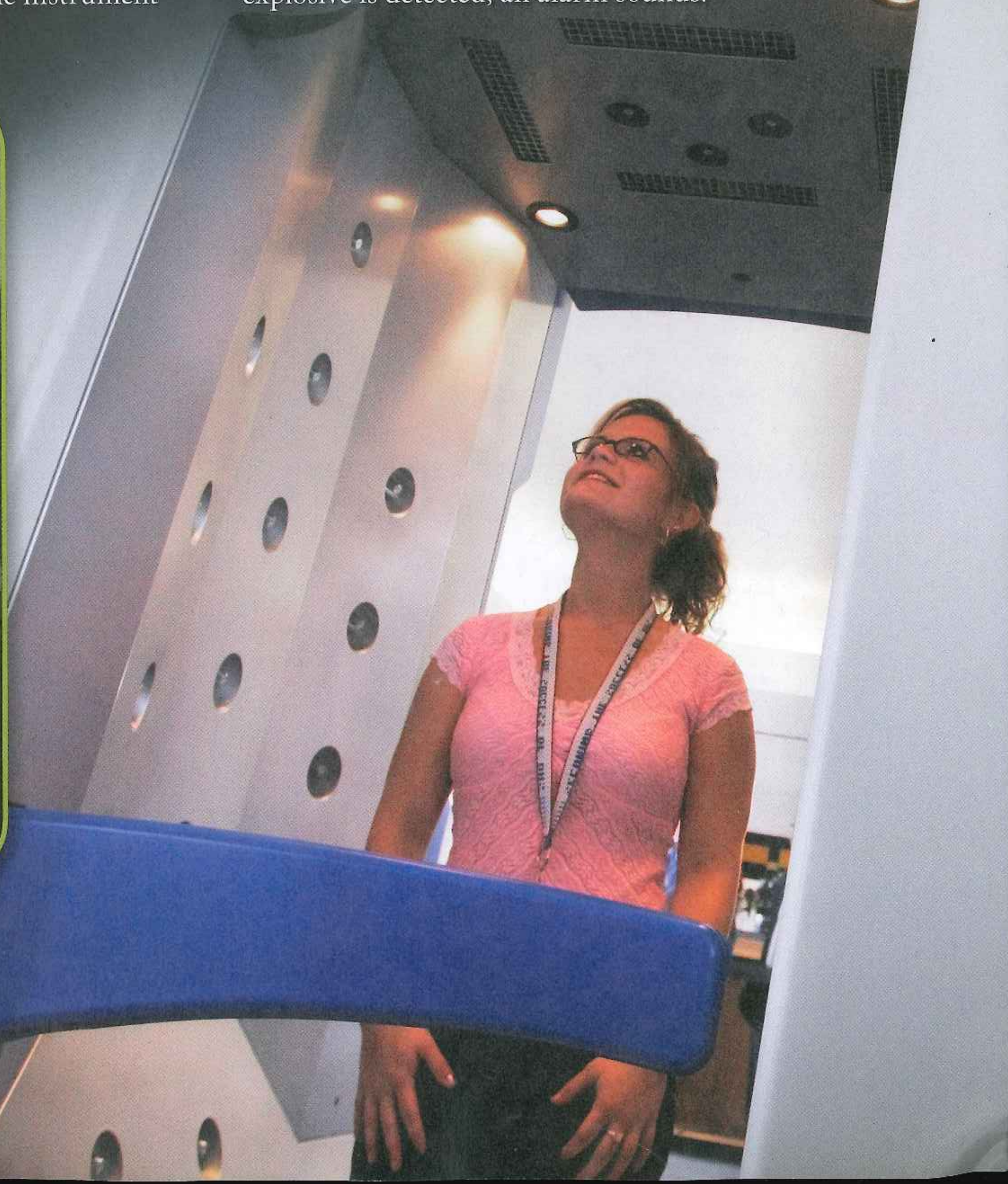
49. The compound methyl butanoate smells like apples. Its percent composition is 58.8% C, 9.8% H, and 31.4% O, and its molar mass is 102 g/mol. What is its empirical formula? What is its molecular formula?

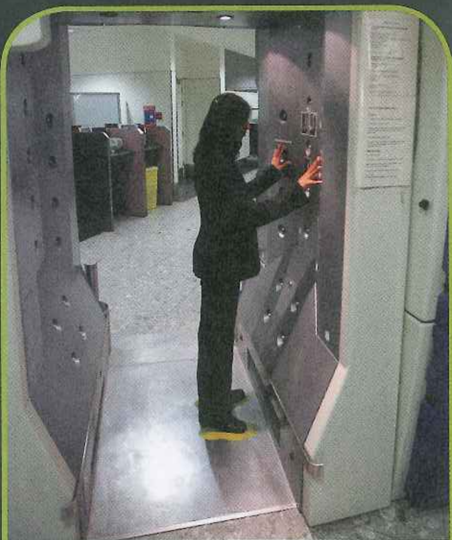
CHEMISTRY & YOU: TECHNOLOGY

Ion Mobility Spectrometry

In 2001, a terrorist boarded an airline flight with explosives inside his shoes. Since that time, Americans have had to remove their shoes during airport security checks. However, newer airport security devices, known as "puffer portals," allow airport security to scan for minute traces of explosives on a person's body and clothing, without the person having to remove any clothing or shoes.

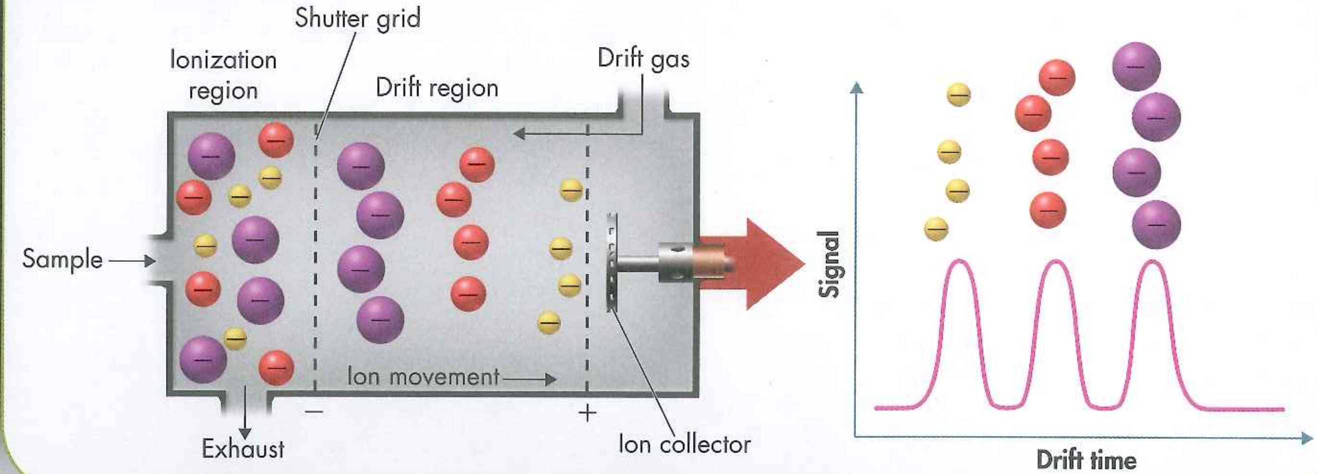
The puffer portal looks like a standard airport metal detector. There are vents and nozzles on the walls and ceiling of the portal. When a passenger steps inside, the doors close, and the instrument sends sharp bursts of air to dislodge particles from his or her body, hair, and clothing. The air sample is then passed through a chemical analysis system called an ion mobility spectrometer (IMS). The IMS identifies compounds based on the amount of time it takes for ions to pass through an electrified field in a tube filled with a nonreactive gas (drift gas). This "drift time" is then compared to a database of drift times of different compounds. In this way, molecules of known explosive or narcotic materials can be detected and identified. If even a picogram of an explosive is detected, an alarm sounds.





PUFFING OUT EXPLOSIVES Bursts of air dislodge particles from a person's hair, body, and clothes. These particles are then directed to an ion mobility spectrometer (IMS).

A Closer Look



IDENTIFYING IONS When the particles enter the IMS, they are ionized, or converted into ions. The ionized particles then travel through a tube containing an electric field, which causes the ions to separate according to their masses, sizes, and shapes. For example, smaller ions move faster and reach the end of the tube before larger ions.

Take It Further

1. Calculate Two common explosive compounds are trinitrotoluene (TNT) and cyclotrimethylenetrinitramine (RDX). The chemical formula of TNT is $C_7H_5N_3O_6$. The chemical formula of RDX is $C_3H_6N_6O_6$. Calculate the molar masses of these two compounds.

2. Analyze Data If TNT and RDX molecules are separated in an IMS solely based on mass, which compound would reach the ion collector first?

3. Predict What do you think would be some other uses for ion mobility spectrometers?

10 Study Guide

BIGIDEA
THE MOLE AND QUANTIFYING MATTER

The mole is an important measurement in chemistry. The mole allows you to convert among the amount of representative particles in a substance, the mass of a substance, and the volume of a gas at STP. The molecular formula of a compound can be determined by first finding the percent composition of the compound and determining the empirical formula. Using the empirical formula mass and the molar mass of the compound, the molecular formula can be determined.

10.1 The Mole: A Measurement of Matter

- Knowing how the count, mass, and volume of an item relate to a common unit allows you to convert among these units.
- The mole allows chemists to count the number of representative particles in a substance.
- The atomic mass of an element expressed in grams is the mass of a mole of the element.
- To calculate the molar mass of a compound, find the number of grams of each element in one mole of the compound. Then add the masses of the elements in the compound.

- mole (308)
- Avogadro's number (308)
- representative particle (308)
- molar mass (313)

10.2 Mole–Mass and Mole–Volume Relationships

- Use the molar mass of an element or compound to convert between the mass of a substance and the moles of the substance.
- The molar volume is used to convert between the number of moles of gas and the volume of the gas at STP.

- Avogadro's hypothesis (320)
- standard temperature and pressure (STP) (320)
- molar volume (320)

10.3 Percent Composition and Chemical Formulas

- The percent by mass of an element in a compound is the number of grams of the element divided by the mass in grams of the compound, multiplied by 100%.
- The percent composition of a compound can be used to calculate the empirical formula of that compound.
- The molecular formula of a compound is either the same as its experimentally determined empirical formula, or it is a simple whole-number multiple of its empirical formula.

- percent composition (325)
- empirical formula (330)

Key Equations

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

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

Math Tune-Up: Mole Problems

Problem	1 Analyze	2 Calculate	3 Evaluate
How many moles of lithium (Li) is 4.81×10^{24} atoms of lithium?	Knowns: number of atoms = 4.81×10^{24} atoms Li 1 mol Li = 6.02×10^{23} atoms Li Unknown: moles = ? mol Li The desired conversion is atoms ⟶ moles.	Use the correct conversion factor to convert from atoms to moles. $4.81 \times 10^{24}\ \cancel{\text{atoms Li}} \times \dfrac{1\ \text{mol Li}}{6.02 \times 10^{23}\ \cancel{\text{atoms Li}}} = \boxed{7.99\ \text{mol Li}}$	The given number of atoms is about 8 times Avogadro's number, so the answer should be around 8 mol of atoms.
Calculate the mass in grams of 0.160 mol H_2O_2.	Known: number of moles = 0.160 mol H_2O_2 Unknown: mass = ? g H_2O_2 The desired conversion is moles ⟶ mass.	Determine the molar mass of H_2O_2 and use the correct conversion factor to convert from moles to grams. $1\ \text{mol}\ H_2O_2 = (2\ \cancel{\text{mol}})(1.0\ \text{g}/\cancel{\text{mol}}) + (2\ \cancel{\text{mol}})(16.0\ \text{g}/\cancel{\text{mol}}) = 34.0\ \text{g}\ H_2O_2$ $0.160\ \cancel{\text{mol}\ H_2O_2} \times \dfrac{34.0\ \text{g}\ H_2O_2}{1\ \cancel{\text{mol}\ H_2O_2}} = \boxed{5.44\ \text{g}\ H_2O_2}$	The number of moles of H_2O_2 is about 0.2, and the molar mass is about 30 g/mol. The answer should be around 6 g.
What is the volume of 1.25 mol He at STP?	Knowns: number of moles = 1.25 mol He 1 mol He at STP = 22.4 L He Unknown: volume = ? L He The desired conversion is moles ⟶ volume at STP.	Use the correct conversion factor to convert from moles to volume at STP. $1.25\ \cancel{\text{mol He}} \times \dfrac{22.4\ \text{L He}}{1\ \cancel{\text{mol He}}} = \boxed{28.0\ \text{L}}$	One mole of gas at STP has a volume of 22.4 L, so 1.25 mol should have a volume larger than 22.4 L.
What is the percent composition of the compound formed when 29.0 g Ag combines completely with 4.30 g S?	Knowns: mass of Ag = 29.0 g Ag mass of S = 4.30 g S mass of compound = 29.0 g + 4.30 g = 33.3 g Unknowns: percent by mass of Ag = ?% Ag percent by mass of S = ?% S Use the equation: % by mass of element = $\dfrac{\text{mass of element}}{\text{mass of compound}} \times 100\%$	Calculate the percent by mass of Ag and S in the compound. $\%\text{Ag} = \dfrac{29.0\ \cancel{\text{g}}}{33.3\ \cancel{\text{g}}} \times 100\%$ $\boxed{\%\text{Ag} = 87.1\%\ \text{Ag}}$ $\%\text{S} = \dfrac{4.30\ \cancel{\text{g}}}{33.3\ \cancel{\text{g}}} \times 100\%$ $\boxed{\%\text{S} = 12.9\%\ \text{S}}$	The percents of the elements add up to 100%.

Hint: Review Sample Problems 10.2 and 10.3 if you have trouble converting between number of representative particles and moles.

Remember: A mole of any gas at STP occupies a volume of 22.4 L.

Hint: Review Sample Problems 10.9 and 10.10 if you have trouble calculating the percent composition of a compound.

10 Assessment

★ Solutions appear in Appendix E

Lesson by Lesson

10.1 The Mole: A Measurement of Matter

50. List three common ways that matter is measured. Give examples of each.

★**51.** Name the representative particle (atom, molecule, or formula unit) of each substance.

a. oxygen gas
b. sodium sulfide
c. sulfur dioxide
d. potassium

★**52.** How many hydrogen atoms are in a representative particle of each substance?

a. $Al(OH)_3$
b. $H_2C_2O_4$
c. $(NH_4)_2HPO_4$
d. $C_4H_{10}O$

53. Describe the relationship between Avogadro's number and one mole of any substance.

★**54.** Find the number of moles in each substance.

a. 2.41×10^{24} formula units of NaCl
b. 9.03×10^{24} atoms of Hg
c. 4.65×10^{24} molecules of NO_2

55. Which contains more molecules: 1.00 mol H_2O_2, 1.00 mol C_2H_6, or 1.00 mol CO?

56. Which contains more atoms: 1.00 mol H_2O_2, 1.00 mol C_2H_6, or 1.00 mol CO?

★**57.** Find the number of representative particles in each substance.

a. 3.00 mol Sn
b. 0.400 mol KCl
c. 7.50 mol SO_2
d. 4.80×10^{-3} mol NaI

58. What is the molar mass of chlorine?

59. List the steps you would take to calculate the molar mass of any compound.

★**60.** Calculate the molar mass of each substance.

a. H_3PO_4
b. N_2O_3
c. $CaCO_3$
d. $(NH_4)_2SO_4$
e. $C_4H_9O_2$
f. Br_2

61. Calculate the mass of 1.00 mol of each of these substances.

a. silicon dioxide (SiO_2)
b. diatomic nitrogen (N_2)
c. iron(III) hydroxide ($Fe(OH)_3$)
d. copper (Cu)

10.2 Mole–Mass and Mole–Volume Relationships

62. Find the mass of each substance.

a. 1.50 mol C_5H_{12}
b. 14.4 mol F_2
c. 0.780 mol $Ca(CN)_2$
d. 7.00 mol H_2O_2
e. 5.60 mol NaOH
f. 3.21×10^{-2} mol Ni

★**63.** Calculate the mass in grams of 0.250 mol of each of the following compounds:

a. sucrose ($C_{12}H_{22}O_{11}$)
b. sodium chloride (NaCl)
c. potassium permanganate ($KMnO_4$)

★**64.** Calculate the number of moles in 1.00×10^2 g of each of the compounds in Problem 63.

65. How many moles is each of the following?

a. 15.5 g SiO_2
b. 0.0688 g AgCl
c. 79.3 g Cl_2
d. 5.96 g KOH
e. 937 g $Ca(C_2H_3O_2)_2$
f. 0.800 g Ca

66. What is the volume of one mole of any gas at STP?

★**67.** Calculate the volume of each of the following gases at STP.

a. 7.64 mol Ar
b. 1.34 mol SO_2
c. 0.442 mol C_2H_6
d. 2.45×10^{-3} mol H_2S

★**68.** A gas has a density of 0.902 g/L at STP. What is the molar mass of this gas?

69. What is the density of each of the following gases at STP?

a. C_3H_8
b. O_2
c. Ne
d. NO_2

70. Find each of the following quantities:

a. the volume, in liters, of 835 g SO_3 at STP
b. the mass, in grams, of a molecule of aspirin ($C_9H_8O_4$)
c. the number of atoms in 5.78 mol NH_4NO_3

10.3 Percent Composition and Chemical Formulas

71. What is the percent composition of the compound formed when 2.70 g of aluminum combine with oxygen to form 5.10 g of aluminum oxide?

*72. Calculate the percent composition when 13.3 g Fe combine completely with 5.7 g O.

*73. Calculate the percent composition of each compound.

a. H_2S
b. $(NH_4)_2C_2O_4$
c. $Mg(OH)_2$
d. Na_3PO_4

*74. Using your answers from Problem 73, calculate the number of grams of these elements.

a. sulfur in 3.54 g H_2S
b. nitrogen in 25.0 g $(NH_4)_2C_2O_4$
c. magnesium in 97.4 g $Mg(OH)_2$
d. phosphorus in 804 g Na_3PO_4

75. Which of the following compounds has the highest percent of iron by mass?

a. $FeCl_2$
b. $Fe(C_2H_3O_2)_3$
c. $Fe(OH)_2$
d. FeO

*76. What is an empirical formula? Which of the following molecular formulas are also empirical formulas?

a. ribose ($C_5H_{10}O_5$)
b. ethyl butyrate ($C_6H_{12}O_2$)
c. chlorophyll ($C_{55}H_{72}MgN_4O_5$)
d. DEET ($C_{12}H_{17}ON$)

77. Which of the following can be classified as an empirical formula?

a. S_2Cl_2
b. $C_6H_{10}O_4$
c. Na_2SO_3

78. Which pair of molecules has the same empirical formula?

a. $C_2H_4O_2$, $C_6H_{12}O_6$
b. $NaCrO_4$, $Na_2Cr_2O_7$

*79. What is the molecular formula for each compound? Each compound's empirical formula and molar mass are given.

a. CH_2O, 90 g/mol
b. HgCl, 472.2 g/mol

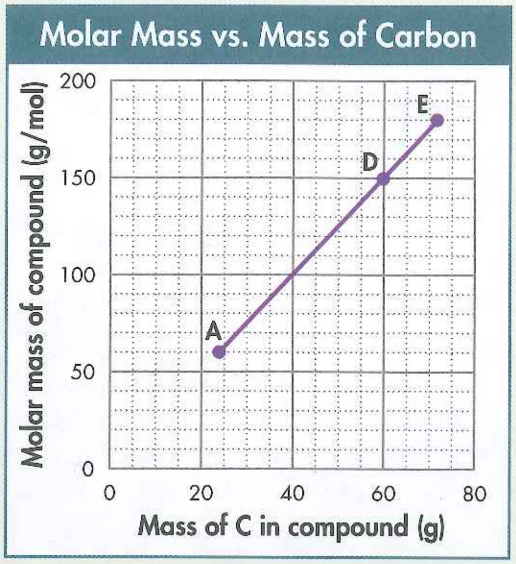
Understand Concepts

*80. Table sugar, or sucrose, has the chemical formula $C_{12}H_{22}O_{11}$.

a. How many atoms are in 1.00 mol of sucrose?
b. How many atoms of C are in 2.00 mol of sucrose?
c. How many atoms of H are in 2.00 mol of sucrose?
d. How many atoms of O are in 3.65 mol of sucrose?

81. How can you determine the molar mass of a gaseous compound if you do not know its molecular formula?

*82. A series of compounds has the empirical formula CH_2O. The graph shows the relationship between the molar mass of the compounds and the mass of carbon in each compound.



a. What are the molecular formulas for the compounds represented by data points A, D, and E?
b. Find the slope of the line. Is this value consistent with the empirical formula? Explain.
c. Two other valid data points fall on the line between points A and D. What are the *x*, *y* values for these data points?

83. Explain what is wrong with each statement.

a. One mole of any substance contains the same number of atoms.
b. A mole and a molecule of a substance are identical in amount.
c. One molar mass of CO_2 contains Avogadro's number of atoms.

84. Which of the following contains the largest number of atoms?

a. 82.0 g Kr
b. 0.842 mol C_2H_4
c. 36.0 g N_2

*85. Calculate the grams of oxygen in 90.0 g of Cl_2O.

86. What is the total mass of a mixture of 3.50×10^{22} formula units Na_2SO_4, 0.500 mol H_2O, and 7.23 g AgCl?

*87. The molecular formula of an antibacterial drug is $C_{17}H_{18}FN_3O_3$. How many fluorine atoms are in a 150-mg tablet of this drug?

88. Determine the empirical formulas of compounds with the following percent compositions:

a. 42.9% C and 57.1% O

b. 32.00% C, 42.66% O, 18.67% N, and 6.67% H

c. 71.72% Cl, 16.16% O, and 12.12% C

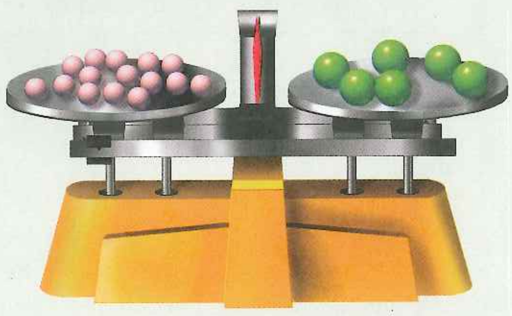
*89. Determine the molecular formula for each compound.

a. 94.1% O and 5.9% H; molar mass = 34 g/mol

b. 50.7% C, 4.2% H, and 45.1% O; molar mass = 142 g/mol

c. 56.6% K, 8.7% C, and 34.7% O; molar mass = 138.2 g/mol

*90. A fictitious "atomic balance" is shown below. Fifteen atoms of boron on the left side of the balance are balanced by six atoms of an unknown element E on the right side.



a. What is the atomic mass of element E?

b. What is the identity of element E?

*91. A typical virus is 5×10^{-6} cm in diameter. If Avogadro's number of these virus particles were laid in a row, how many kilometers long would the line be?

92. Calculate the empirical formula for each compound.

a. compound consisting of 0.40 mol Cu and 0.80 mol Br

b. compound with 4 atoms of carbon for every 12 atoms of hydrogen

93. Muscle fatigue can result from the buildup of lactic acid. The percent composition of lactic acid is 40.0% C, 6.67% H, and 53.3% O. What is the molecular formula of lactic acid if its molar mass is 90.0 g/mol?

*94. What mass of helium is needed to inflate a balloon to a volume of 5.50 L at STP?

95. How many water molecules are in a 1.00-L bottle of water? The density of water is 1.00 g/mL.

Think Critically

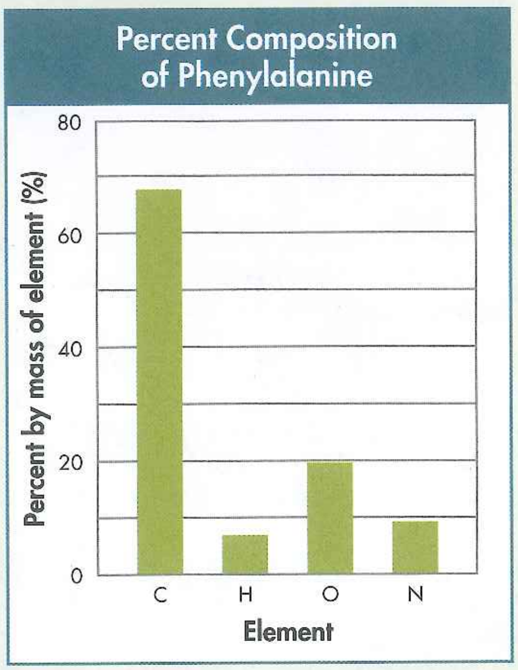
*96. **Infer** What is the empirical formula of a compound that has three times as many hydrogen atoms as carbon atoms but only half as many oxygen atoms as carbon atoms?

97. **Apply Concepts** How are the empirical and molecular formulas of a compound related?

98. **Compare** Why does one mole of carbon have a smaller mass than one mole of sulfur? How are the atomic structures of these elements different?

99. **Analyze Data** One mole of any gas at STP equals 22.4 L of that gas. It is also true that different elements have different atomic volumes, or diameters. How can you reconcile these two statements?

*100. **Interpret Graphs** The graph shows the percent composition of phenylalanine.



a. What is the empirical formula for phenylalanine?

b. If the molar mass of phenylalanine is 165.2 g/mol, what is its molecular formula?

Enrichment

*101. **Infer** Nitroglycerine contains 60% as many carbon atoms as hydrogen atoms, three times as many oxygen atoms as nitrogen atoms, and the same number of carbon and nitrogen atoms. The number of moles of nitroglycerine in 1.00 g is 0.00441. What is the molecular formula of nitroglycerine?

102. **Calculate** The density of nickel is 8.91 g/cm^3. How large a cube, in cm^3, would contain 2.00×10^{24} atoms of nickel?

*103. **Calculate** Dry air is about 20.95% oxygen by volume. Assuming STP, how many oxygen molecules are in a 75.0-g sample of air? The density of air is 1.19 g/L.

104. **Graph** The table below gives the molar mass and density of seven gases at STP.

Substance	Molar mass (g/mol)	Density (g/L)
Oxygen	32.0	1.43
Carbon dioxide	44.0	1.96
Ethane	30.0	1.34
Hydrogen	2.0	0.089
Sulfur dioxide	64.1	2.86
Ammonia	17.0	0.759
Fluorine	38.0	1.70

a. Plot these data, with density on the *x*-axis.
b. What is the slope of the straight-line plot?
c. What is the molar mass of a gas at STP that has a density of 1.10 g/L?
d. A mole of a gas at STP has a mass of 56.0 g. Use the graph to determine its density.

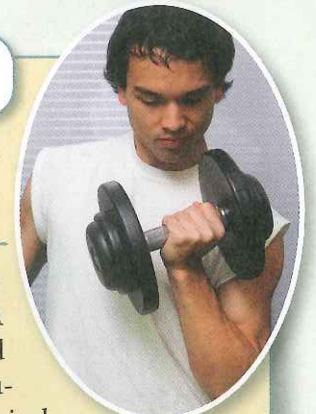
*105. **Calculate** Avogadro's number has been determined by about 20 different methods. In one approach, the spacing between ions in an ionic substance is determined by using a technique called X-ray diffraction. X-ray diffraction studies of sodium chloride have shown that the distance between adjacent Na^+ and Cl^- ions is 2.819×10^{-8} cm. The density of solid NaCl is 2.165 g/cm^3. By calculating the molar mass to four significant figures, you can determine Avogadro's number. What value do you obtain?

106. **Use Models** In Chapter 3, you learned that the densities of solids and liquids are measured in g/cm^3, but the densities of gases are measured in g/L. Draw atomic diagrams of a solid and of a gas that show why the two different units are practical.

Write About Science

107. **Connect to the BIGIDEA** Research the history of Avogadro's number. What elements other than carbon have been used to define a mole? Write a report that summarizes your findings.

CHEMYSTERY



A Formula for Cheating

Typically, steroids can be detected in an athlete's urine. A urine sample is collected and is first injected into an instrument that separates the chemical compounds in the urine.

The separated compounds are then analyzed using a mass spectrometer. The mass spectrometer provides information such as the molar mass of the compounds present in the urine sample and the molecular structure of these compounds. These structures can be compared against a database of known compounds to identify the presence of steroids in the sample.

*108. **Calculate** Analysis of an athlete's urine found the presence of a compound with a molar mass of 312 g/mol. How many moles of this compound are contained in 30.0 mg? How many molecules of the compound is this?

109. **Connect to the BIGIDEA** The compound found in the athlete's urine, the steroid THG, has a percent composition of 80.8% carbon, 8.97% hydrogen, and 10.3% oxygen. What is the empirical formula of THG? If the molar mass of THG is 312 g/mol, what is the molecular formula?

Cumulative Review

*110. Identify at least one chemical change and two physical changes that are occurring in the photo.

*111. Classify each of the following as a physical change or a chemical change.

a. An aspirin tablet is crushed to a powder.
b. A red rose turns brown.
c. Grape juice turns to wine.
d. Fingernail polish remover evaporates.
e. A bean seed sprouts.
f. A piece of copper is beaten into a thin sheet.

112. Which of these statements are true about every solution?

a. Solutions are in the liquid state.
b. Solutions are homogeneous.
c. Solutions are mixtures.
d. Solutions are composed of at least two compounds.

113. A student writes down the density of table sugar as 1.59 and the density of carbon dioxide as 1.83. Can these values be correct? Explain.

*114. A block of wood measuring 2.75 cm $\times$ 4.80 cm $\times$ 7.50 cm has a mass of 84.0 g. Will the block of wood sink or float in water?

*115. Convert each of the following:

a. 4.72 g to mg
b. 2.7×10^3 cm/s to km/h
c. 4.4 mm to dm

*116. How many protons, electrons, and neutrons are in each isotope?

a. zirconium-90
b. palladium-108
c. bromine-81
d. antimony-123

*117. Write the complete electron configuration for each atom.

a. fluorine
b. lithium
c. rubidium

118. Why do the elements magnesium and barium have similar chemical and physical properties?

*119. Which of the following are transition metals: Cr, Cd, Ca, Cu, Co, Cs, Ce?

120. How can the periodic table be used to infer the number of valence electrons in an atom?

121. How does a molecule differ from an atom?

122. Draw electron dot structures and predict the shapes of the following molecules:

a. PH_3
b. CO
c. CS_2
d. CF_4

123. How are single, double, and triple bonds indicated in electron dot structures?

124. Give an example of each of the following:

a. coordinate covalent bonding
b. resonance structures
c. exceptions to the octet rule

125. Explain how you can use electronegativity values to classify a bond as nonpolar covalent, polar covalent, or ionic.

*126. Identify any incorrect formulas among the following:

a. H_2O_2
b. $NaIO_4$
c. SrO
d. CaS_2
e. $CaHPO_4$
f. BaOH

*127. Name these compounds.

a. $Fe(OH)_3$
b. NH_4I
c. Na_2CO_3
d. CCl_4

*128. Write formulas for these compounds.

a. potassium nitrate
b. copper(II) oxide
c. magnesium nitride
d. silver fluoride

If You Have Trouble With . . .

Question	110	111	112	113	114	115	116	117	118	119	120	121	122	123	124	125	126	127	128
See Chapter	2	2	2	3	3	3	4	5	6	6	7	8	8	8	8	8	9	9	9

Standardized Test Prep

Tips for Success

Wear a Watch Be aware of how many questions you have to answer and how much time you have to answer them. Look at your watch or a clock frequently to keep track of your progress.

1. Choose the term that best completes the second relationship.

a. dozen : eggs
mole : _____

(A) atoms (C) size
(B) 6.02×10^{23} (D) grams

b. mole : Avogadro's number
molar volume : _____

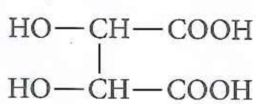
(A) mole (C) STP
(B) water (D) 22.4 L

Select the choice that best answers each question or completes each statement.

2. Calculate the molar mass of ammonium phosphate, $(NH_4)_3PO_4$.

(A) 113.0 g/mol (C) 149.0 g/mol
(B) 121.0 g/mol (D) 242.0 g/mol

3. Based on the structural formula below, what is the empirical formula for tartaric acid, a compound found in grape juice?



(A) $C_2H_3O_3$ (C) CHO
(B) $C_4H_6O_6$ (D) $C_1H_{1.5}O_{1.5}$

4. How many hydrogen atoms are in six molecules of ethylene glycol, $C_2H_6O_2$?

(A) 6 (C) $6 \times (6.02 \times 10^{23})$
(B) 36 (D) $36 \times (6.02 \times 10^{23})$

5. Which of these compounds has the largest percent by mass of nitrogen?

(A) N_2O (D) N_2O_3
(B) NO (E) N_2O_4
(C) NO_2

6. Which of these statements is true of a balloon filled with 1.00 mol $N_2(g)$ at STP?

I. The balloon has a volume of 22.4 L.
II. The contents of the balloon have a mass of 14.0 g.
III. The balloon contains 6.02×10^{23} molecules.

(A) I only (C) I and III only
(B) I and II only (D) I, II, and III

7. Allicin, $C_6H_{10}S_2O$, is the compound that gives garlic its odor. A sample of allicin contains 3.0×10^{21} atoms of carbon. How many hydrogen atoms does this sample contain?

(A) 10 (C) 1.8×10^{21}
(B) 1.0×10^{21} (D) 5.0×10^{21}

The lettered choices below refer to Questions 8–11. A lettered choice may be used once, more than once, or not at all.

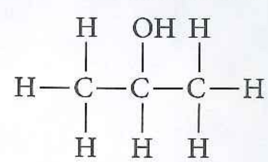
(A) CH (B) CH_2 (C) C_2H_5 (D) CH_3 (E) C_2H_3

Which of the formulas is the empirical formula for each of the following compounds?

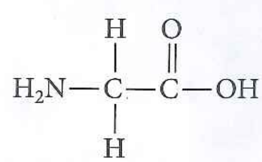
8. C_8H_{12} **10.** C_2H_6
9. C_6H_6 **11.** C_4H_{10}

For Questions 12–14, write the molecular formula for each compound whose structural formula is shown. Then calculate the compound's molar mass.

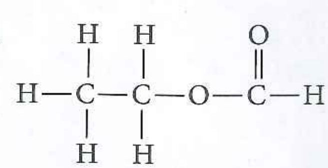
12.



13.



14.



If You Have Trouble With . . .

Question	1	2	3	4	5	6	7	8	9	10	11	12	13	14
See Lesson	10.2	10.1	10.3	10.1	10.3	10.2	10.1	10.3	10.3	10.3	10.3	10.1	10.1	10.1