10.2 Mole–Mass and Mole–Volume Relationships

Connecting to Your World Guess how many jelly beans are in the container and win a prize! You decide to enter the contest and you win. Was it just a lucky guess? Not exactly. You estimated the length and diameter of a jelly bean 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 chemistry problems. In this section you will find out how the mole and mass are related.

Guide for Reading

Key Concepts

- How do you convert the mass of a substance to the number of moles of the substance?
- What is the volume of a gas at STP?

Vocabulary

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

Reading Strategy

Monitoring Your Understanding Before you read, preview the key concepts, the section heads, the boldfaced terms, and the visuals. List three things you expect to learn. After reading, state what you learned about each item you listed.

The Mole–Mass Relationship

In the previous section, 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 what the molar mass of oxygen is? 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 (2×16.0 g). 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. You can avoid confusion such as this by using the formula of the substance, in this case, O_2 or O.

Suppose you need 3.00 mol of sodium chloride (NaCl) for a laboratory experiment. How can you measure this amount? It would be convenient to use a balance to measure the mass. But what mass in grams is 3.00 mol of NaCl? **Use the molar mass of an element or compound to convert between the mass of a substance and the moles of a substance.** The conversion factor for the calculation is based on the relationship: molar mass = 1 mol. Use the following equation to calculate the mass in grams of a given number of moles.

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

The molar mass of NaCl is 58.5 g/mol, so the mass of 3.00 mol NaCl is calculated in this way.

$$\text{mass of NaCl} = 3.00\ \cancel{\text{mol}} \times \frac{58.5\text{ g}}{1\ \cancel{\text{mol}}} = 176\text{ g}$$

When you measure 176 g of NaCl on a balance, you are measuring 3.00 moles of NaCl.

Section Resources

Print

- **Guided Reading and Study Workbook,** Section 10.2
- **Core Teaching Resources,** Section 10.2 Review
- **Laboratory Manual,** Lab 12
- **Small-Scale Chemistry Laboratory Manual,** Lab 13
- **Transparencies,** T108–T109

Technology

- **Virtual Chemistry Labs,** Lab 3
- **Interactive Textbook with ChemASAP,** Problem-Solving 10.16, 10.18, 10.20, 10.22; Simulation 10; Assessment 10.2

10.2

1 FOCUS

Objectives

10.2.1 Describe how to convert the mass of a substance to the number of moles of a substance, and moles to mass.

10.2.2 Identify the volume of a quantity of gas at STP.

Guide for Reading

Build Vocabulary L2

Graphic Organizer Have students divide a piece of paper into three columns. In the first column, have students list what they know about each vocabulary term. In the second column, have them list what they want to know about each term. As they progress through the chapter, have them list in the third column what they learn about each term.

Reading Strategy L2

Outline As students read this section, have them outline the concepts, using the section headings as the headings in the outline.

2 INSTRUCT

Connecting to Your World

Ask, **What method, other than estimating the volume of each individual candy, might you use to determine the number of jelly beans present?** *(Answers might include using the mass of one jelly bean and the total mass of all the jelly beans.)*

The Mole–Mass Relationship

Discuss L2

Review the mathematical conversions of moles to number of particles and number of particles to moles. Stress that using dimensional analysis in problem solving allows students to perform these calculations without having to memorize the process. Emphasize the use of units when solving problems.

Sample Problem 10.5

Answers

16. 4.52×10^{-3} mol $C_{20}H_{42} \times 282.0$ g $C_{20}H_{42}$/1 mol $C_{20}H_{42} = 1.27$ g $C_{20}H_{42}$

17. 2.50 mol $Fe(OH)_2 \times 89.8$ g $Fe(OH)_2$/1 mol $Fe(OH)_2 = 225$ g $Fe(OH)_2$

Practice Problems Plus L2

Calculate the mass in grams for 0.250 mol of each of the following compounds:

a. sucrose *(85.5 g)*
b. sodium chloride *(14.6 g)*
c. potassium permanganate *(39.5 g)*

Math Handbook

For a math refresher and practice, direct students to significant figures, page R59.

Relate L2

In chemical manufacturing processes, reactants are purchased by mass, and products are sold by mass. However, each process is set up based on the knowledge of the ratio in which moles of reactants combine with each other to form moles of products. Ask, **If one mole of reactant produces one mole of product, how can you use the information on pp. 298 and 299 to find the mass of product if you know the mass of the reactant?** *(Change the mass of reactant to number of moles. Then, change that number of moles to mass of product.)* Tell students that they will learn more about conversions such as this in Chapter 12.

CLASS Activity

Problem Solving L2

Students often can do one type of problem but have difficulty when two types of problems are interspersed. Have students do some practice problems for Sample Problems 10.5 and 10.6 individually. Then, have them do a mixture of the two types of problems.

Math Handbook

For help with significant figures go to page R59.

Interactive Textbook

Problem-Solving 10.16
Solve Problem 16 with the help of an interactive guided tutorial.

with ChemASAP

SAMPLE PROBLEM 10.5

Converting Moles to Mass

The aluminum satellite dishes in Figure 10.8 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 of 9.45 mol of aluminum oxide?

1 Analyze ***List the known and the unknown.***

Known
- number of moles = 9.45 mol Al_2O_3

Unknown
- mass = ? g Al_2O_3

The mass of the compound is calculated from the known number of moles of the compound. The desired conversion is moles ⟶ mass.

2 Calculate ***Solve for the unknown.***

Determine the molar mass of Al_2O_3: 1 mol Al_2O_3 = 102.0 g Al_2O_3
Multiply the given number of moles by the conversion factor relating moles of Al_2O_3 to grams of Al_2O_3.

$$\text{mass} = 9.45\ \cancel{\text{mol Al}_2\text{O}_3} \times \frac{102.0\ \text{g Al}_2\text{O}_3}{1\ \cancel{\text{mol Al}_2\text{O}_3}}$$
$$= 964\text{g Al}_2\text{O}_3$$

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 about 1000 g. The answer has been rounded to the correct number of significant figures.

Practice Problems

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.

Figure 10.8 These aluminum satellite dishes at the National Radio Astronomy Observatory near Socorro, New Mexico are naturally protected from corrosion by the formation of a thin film of aluminum oxide (Al_2O_3).

In Sample Problem 10.5, you used a conversion factor based on the molar mass to convert moles to mass. Now suppose that in a laboratory experiment you obtain 10.0 g of sodium sulfate (Na_2SO_4). How many moles is this? You can calculate the number of moles using the same relationship you used in Sample Problem 10.5, 1 mol = molar mass, but this time the conversion factor is inverted. Use the following equation to convert your 10.0 g of Na_2SO_4 into moles.

$$\text{moles} = \text{mass (grams)} \times \frac{1\ \text{mole}}{\text{mass (grams)}}$$

The molar mass of Na_2SO_4 is 142.1 g/mol, so the number of moles of Na_2SO_4 is calculated this way.

$$\text{moles of } Na_2SO_4 = 10.0\ \text{g} \times \frac{1\ \text{mol}}{142.1\ \text{g}} = 7.04 \times 10^{-2}\ \text{mol}$$

Checkpoint ***What conversion factor should you use to convert mass to moles?***

Rust weakens an iron chain.

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.***

Known	Unknown
• mass = 92.2 g Fe_2O_3	• number of moles = ? mol Fe_2O_3

The unknown number of moles of the compound is calculated from a known mass of a compound. The conversion is mass $\longrightarrow$ moles.

2 Calculate ***Solve for the unknown.***

Determine the molar mass of Fe_2O_3: 1 mol = 159.6 g Fe_2O_3

Multiply the given mass by the conversion factor relating mass of Fe_2O_3 to moles of Fe_2O_3.

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

$$= 0.578\ \text{mol}\ Fe_2O_3$$

3 Evaluate ***Does the result make sense?***

Because the given mass (about 90 g) is slightly larger than the mass of one-half mole of Fe_2O_3 (about 160 g), the answer should be slightly larger than one-half (0.5) mol.

Practice Problems

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.

Math Handbook

For help with using a calculator, go to page R62.

Interactive Textbook

Problem-Solving 10.18

Solve Problem 18 with the help of an interactive guided tutorial.

with ChemASAP

Sample Problem 10.6

Answers

18. 3.70×10^{-1} g B $\times$ 1 mol B/10.8 g B = 3.43×10^{-2} mol B

19. 75.0 g N_2O_3 $\times$ 1 mol N_2O_3/76.0 g N_2O_3 = 0.987 mol N_2O_3

Practice Problems Plus L2

Calculate the number of moles in 1.00×10^2 g of each of the following compounds.

a. sucrose *(0.292 mol)*

b. sodium chloride *(1.71 mol)*

c. potassium permanganate *(0.633 mol)*

Math Handbook

For a math refresher and practice, direct students to using a calculator, page R62.

Differentiated Instruction

Less Proficient Readers L1

Provide students with several problems and have them analyze each problem for what information it contains, that is, help students identify the known(s) and the unknown(s). Also have them look for key words, such as *total, difference,* and *larger,* that show operations or relationships.

Answers to...

Checkpoint 1 mole/mass (grams)

The Mole–Volume Relationship

Discuss L2

Ask, **What unit is used for the mass of a mole?** *(grams per mole, g/mol)* **What unit is used for the volume of a mole?** *(L per mole, L/mol)* Point out to students that, unlike solids and liquids, the molar volume of gases is predictable but is affected by temperature and pressure. Ask, **How does temperature affect the volume of a gas?** *(When temperature increases, volume increases. When temperature decreases, volume decreases.)* **How does pressure affect the volume of a gas?** *(When pressure increases, volume decreases. A decrease in pressure causes an increase in volume.)* Emphasize that when comparing the molar volumes of gases, it is necessary to have the gases at the same conditions of temperature and pressure.

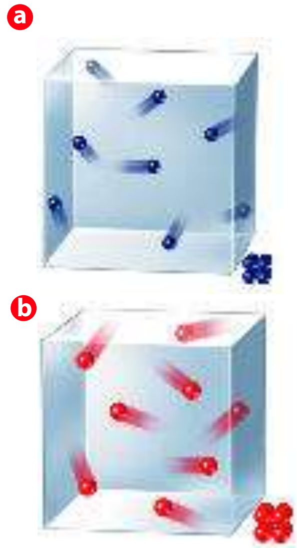

Figure 10.9 In each container, the volume occupied by the gas molecules is small compared with the container's volume, so the molecules are not tightly packed. **a** The molecules in this container are small. **b** This container can accommodate the same number of larger molecules.

The Mole–Volume Relationship

Look back at Figure 10.7. 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 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 should this be?

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. But 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 10.9.

If you buy a party balloon filled with helium and take it home on a cold day, you might notice that the balloon shrinks while it is outside. The volume of a gas varies with a change in temperature. The volume of a gas also varies with a change in pressure. In Figure 10.10, notice the changes in an empty water bottle when it is in the cabin of an airplane while in flight and after the plane has landed. The trapped air occupies the full volume of the bottle in the cabin where the air pressure is lower than it is on the ground. The increase in pressure when the plane lands causes the volume of the air in the bottle to decrease. Because of these variations due to 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.** Figure 10.11 gives you an idea of the size of 22.4 L. The quantity, 22.4 L, is called the **molar volume** of a gas.

Checkpoint ***What is meant by standard temperature and pressure?***

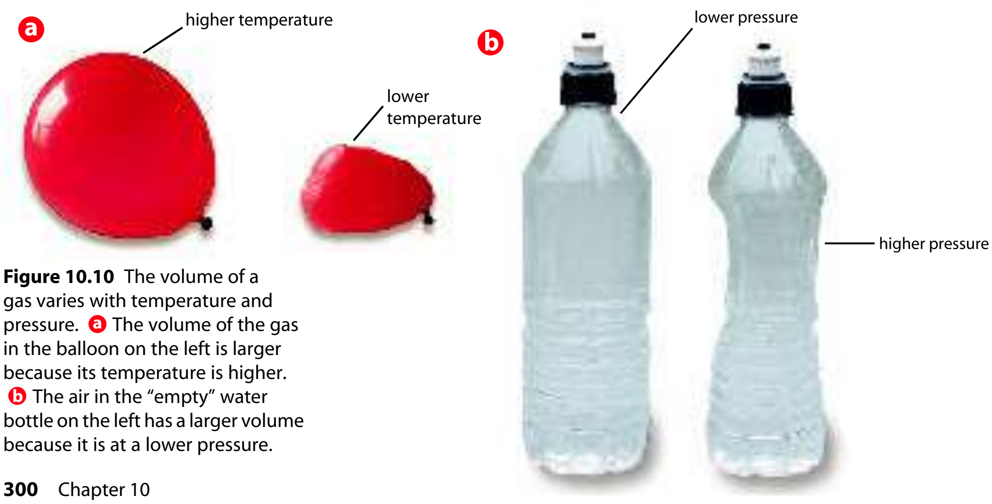

Figure 10.10 The volume of a gas varies with temperature and pressure. **a** The volume of the gas in the balloon on the left is larger because its temperature is higher. **b** The air in the "empty" water bottle on the left has a larger volume because it is at a lower pressure.

Differentiated Instruction

Gifted and Talented L3

Have students apply their problem-solving skills to this question: Students heated a mixture of potassium chlorate and manganese dioxide, producing 0.377 L of oxygen gas at STP. **What was the mass of the gas collected?** *(0.539 g)*

300 Chapter 10

Calculating Volume at STP The molar volume is used to convert a known number of moles of gas to the volume of the gas at STP. The relationship 22.4 L = 1 mol at STP provides the conversion factor.

$$\text{volume of gas} = \text{moles of gas} \times \frac{22.4\ \text{L}}{1\ \text{mol}}$$

Suppose you have 0.375 mol of oxygen gas and want to know what volume the gas will occupy at STP.

$$\text{volume of } O_2 = 0.375\ \cancel{\text{mol}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}} = 8.40\ \text{L}$$

SAMPLE PROBLEM 10.7

Calculating the Volume of a Gas 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.***

Knowns
- moles = 0.60 mol SO_2
- 1 mol SO_2 = 22.4 L SO_2

Unknown
- volume = ? L SO_2

Use the relationship 1 mol SO_2 = 22.4 L SO_2 (at STP) to write the conversion factor needed to convert moles to volume.

The conversion factor is $\frac{22.4\ \text{L } SO_2}{1\ \text{mol } SO_2}$.

2 Calculate ***Solve for the unknown.***

$$\text{volume} = 0.60\ \cancel{\text{mol } SO_2} \times \frac{22.4\ \text{L } SO_2}{1\ \cancel{\text{mol } SO_2}} = 13\ \text{L } SO_2$$

3 Evaluate ***Does the result make sense?***

Because 1 mol of any gas at STP has a volume of 22.4 L, 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.

Practice Problems

20. What is the volume of these gases at STP?
a. 3.20×10^{-3} mol CO_2
b. 3.70 mol N_2

21. At STP, what volume do these gases occupy?
a. 1.25 mol He
b. 0.335 mol C_2H_6

The opposite conversion, from the volume of a gas at STP to the number of moles of gas, uses the same relationship: 22.4 L = 1 mol at STP. Suppose, in an experiment, you collect 0.200 liter of hydrogen gas at STP. You can calculate the number of moles of hydrogen in this way.

$$\text{moles} = 0.200\ \cancel{\text{L } H_2} \times \frac{1\ \text{mol } H_2}{22.4\ \cancel{\text{L } H_2}} = 8.93 \times 10^{-3}\ \text{mol } H_2$$

Figure 10.11 This box, with a volume of 22.4 L, holds one mole of gas at STP.

Math Handbook
For help with dimensional analysis, go to page R66.

Interactive Textbook
Problem-Solving 10.20
Solve Problem 20 with the help of an interactive guided tutorial.
with ChemASAP

Sample Problem 10.7

Answers

20. a. 3.20×10^{-3} mol $CO_2 \times 22.4$ L CO_2/1 mol $CO_2 = 7.17 \times 10^{-2}$ L CO_2
b. 3.70 mol $N_2 \times 22.4$ L N_2/1 mol N_2 = 82.9 L N_2

21. a. 1.25 mol He × 22.4 L He/1 mol He = 28.0 L He
b. 0.335 mol $C_2H_6 \times 22.4$ L C_2H_6/1 mol C_2H_6 = 7.50 L C_2H_6

Practice Problems Plus L2

At STP, what volume is occupied by each of the following gases?
a. 1.34 mol SO_2 *(30.0 L SO_2)*
b. 2.45×10^{-3} mol H_2S *(5.49×10^{-2} L H_2S)*
c. 6.7 mol H_2 *(1.5×10^2 L H_2)*

Math Handbook
For a math refresher and practice, direct students to dimensional analysis, page R66.

TEACHER Demo

Molar Volume L2

Purpose Students observe an approximation of molar volume.

Materials Dry ice, towel, hammer, large plastic bag, duct tape, tongs, beaker, balance

Procedure Wrap the dry ice in a towel and hit it with the hammer until it is in small pieces. Place 44 g (1 mol CO_2) of the small pieces in a beaker. Expel any air from the plastic bag, and tape the opening of the bag securely over the top of the beaker. As the dry ice sublimes, the bag will inflate.

Safety Wear goggles while crushing the dry ice, and do not allow dry ice to contact skin. Use tongs to handle the dry ice.

Expected Outcomes The volume of gas produced will not equal 22.4 L because conditions are not standard. However, the volume will be close to this value.

Differentiated Instruction

Gifted and Talented L3

Have students design a way to measure the volume of CO_2 produced in the Teacher Demo on this page. Methods might include measuring how much water the filled bag displaces.

Answers to...

Checkpoint 0°C and 101.3 KPa or 1 atmosphere (atm)

Discuss L2

Review the concept of density as a ratio of mass to volume. Discuss the units that go with density. *(g/mL, g/cm^3, or g/L)* Ask, **If you had a mole of gas at STP, how could you calculate the density?** *(molar mass/22.4 L = density)*

Sample Problem 10.8

Answers

22. 3.58 g/1 L × 22.4 L/1 mol = 80.2 g/mol

23. 83.8 g/1 mol × 1 mol/22.4 L = 3.74 g/L

Practice Problems Plus L2

1. A gas has a density of 0.902 g/L. What is the molar mass of this gas? *(20.2 g/mol)*

2. What is the density of oxygen gas at STP? *(1.43 g/L)*

Math Handbook

For a math refresher and practice, direct students to significant figures, page R59.

Discuss L2

Emphasize to students that if the number of moles is known, the mass of the substance or the volume of a gas can be calculated. This concept will continue to be essential as students study mass–mass and other stoichiometric relationships in Chapter 12.

Calculating Molar Mass from Density A gas-filled balloon will either sink or float in the air depending on whether the density of the balloon's gas 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.

$$\text{molar mass} = \text{density at STP} \times \text{molar volume at STP}$$

$$\frac{\text{grams}}{\text{mole}} = \frac{\text{grams}}{\text{L}} \times \frac{22.4\ \text{L}}{1\ \text{mole}}$$

Checkpoint ***How is the density of a gas usually measured?***

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?

1 Analyze ***List the knowns and the unknown.***

Knowns	Unknown
• density = 1.964 g/L	• molar mass = ? g/mol
• 1 mol (gas at STP) = 22.4 L	

The conversion factor needed to convert density to molar mass is $\frac{22.4\ \text{L}}{1\ \text{mol}}$.

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

2 Calculate ***Solve for the unknown.***

$$\text{molar mass} = \frac{1.964\ \text{g}}{1\ \cancel{\text{L}}} \times \frac{22.4\ \cancel{\text{L}}}{1\ \text{mol}}$$
$$= 44.0\ \text{g/mol}$$

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.

Practice Problems

22. A gaseous compound composed of sulfur and oxygen, which is linked to the formation of acid rain, has a density of 3.58 g/L at STP. What is the molar mass of this gas?

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

Math Handbook

For help with significant figures, go to page R59.

Interactive Textbook

Problem-Solving 10.22
Solve Problem 22 with the help of an interactive guided tutorial.

with ChemASAP

Differentiated Instruction

Special Needs L1

To help students learn to solve conversion problems, have them write each of the three pairs of conversion factors in Figure 10.12 on the two sides of a 3 × 5 card. For example, have them write

$$\frac{\text{molar mass (grams)}}{1.00\text{mol}}$$

on one side of a card and

$$\frac{1.00\text{mol}}{\text{molar mass (grams)}}$$

on the other. Also have them make a card with each of the units of the known and unknown quantities: *grams, moles, liters,* and *representative particles*. Students can first set up a problem using the cards in a way that will produce the desired unit for the answer and then fill in the actual numbers to solve the problem.

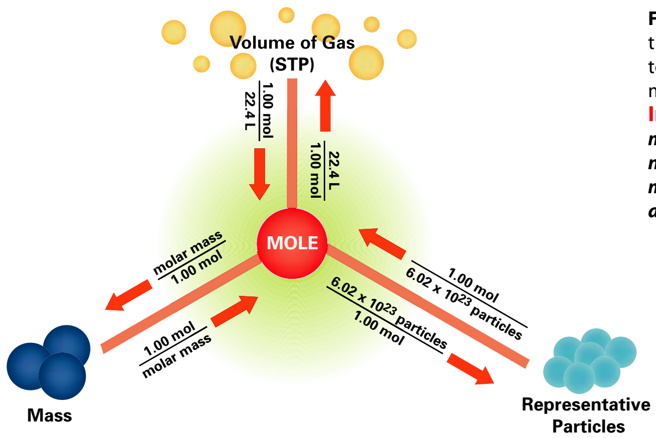

Figure 10.12 The map shows the conversion factors needed to convert among volume, mass, and number of particles. **Interpreting Diagrams** ***How many conversion factors are needed to convert from the mass of a gas to the volume of a gas at STP?***

The Mole Road Map

You have now examined a mole in terms of particles, mass, and volume of gases at STP. Figure 10.12 summarizes these relationships and illustrates the importance of the mole. 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.

Interactive Textbook

Simulation 10 Use the mole road map to convert among mass, volume, and number of representative particles.

with ChemASAP

10.2 Section Assessment

24. **Key Concept** Describe how to convert between the mass and the number of moles of a substance.

25. **Key Concept** What is the volume of one mole of any gas at STP?

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

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

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

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

30. 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).

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

Connecting Concepts

Density 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 a gas that show why the two different units are practical.

Interactive Textbook

Assessment 10.2 Test yourself on the concepts in Section 10.2.

with ChemASAP

Section 10.2 Assessment

24. To convert mass to moles, multiply the given mass by 1 mol/molar mass. To convert moles to mass, multiply the given number of moles by molar mass/1 mol.

25. 22.4 L

26. 567 g $CaCO_3$

27. 11.0 mol C_2H_6O

28. 33.6 L Cl_2

29. 39.9 g/mol

30. gas A: 28.0 g, nitrogen
gas B: 64.1 g, sulfur dioxide
gas C: 16.0 g, methane

31. The balloons have the same number of molecules. Each balloon is filled with one mole of gas, and one mole of any gas has the same number of molecules. The masses of the balloons will differ.

The Mole Road Map

Use Visuals L1

Figure 10.12 Have students study the figure. Guide them through examples of the various mole conversions. For example, start with 50.0 g of a compound or element and convert it to moles, and then to particles. Then, start with a certain volume of a gas and convert it to mass or particles.

3 ASSESS

Evaluate Understanding L2

Have students work problems in which they use molar mass and molar volume to calculate the densities of gases. Have volunteers come to the board to show their calculations. Have the class determine whether the calculations are correct.

Reteach L1

Review the concept of density as a ratio of mass to volume. Ask, **If you know the molar volume of a gas, how could density help you determine the molar mass?** *(Molar mass is the product of density and molar volume.)*

Connecting Concepts

Student answers should show diagrams of particles that are close enough to touch each other for liquids and solids. The diagrams of gases should show the particles far apart.

Interactive Textbook

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 10.2.

with ChemASAP

Answers to...

Figure 10.12 2

Checkpoint The density of a gas is measured in grams per liter (g/L).

Small-Scale LAB

Counting by Measuring Mass L2

Objective After completing this activity, students will be able to:

- measure masses of chemicals and convert their data to moles and atoms.
- explore the quantitative chemical compositions of common substances

Skills Focus measuring, calculating

Prep Time 20 minutes
Class Time 40 minutes

Teaching Tips

- Explore with students ways of finding the mass of liquid and solid samples so that the container does not interfere with the measurement.
- If time allows, have students repeat the procedure and average the data.

Expected Outcome

See Data Table.

Analyze

Sample calculations using sample data:

1. 5.09 g NaCl × 1 mol NaCl/58.5 g NaCl = 0.0870 mol NaCl
2. See data table for answers.
3. See data table for answers.
4. 0.478 mol H × 6.02×10^{23} atoms/1 mol H = 2.88×10^{23} atoms H
5. water
6. water

You're the Chemist!

Sample answers are provided.

1. Find the mass of 100 drops of water, and then calculate the mass in grams per drop.
2. Find the mass of a piece of chalk. Write your name and find the mass of the chalk again. Convert the mass difference to moles and atoms.

For Enrichment L3

Have students use their results from the lab to calculate the volume of one mole of each substance tested. Then, have them use a balance to measure one mole of each substance and a graduated cylinder to find its volume. Have them compare the calculated and experimental values and discuss any discrepancies.

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

- chemicals shown in the table
- plastic spoon
- weighing paper
- watchglass or small beaker
- balance
- paper
- pencil
- ruler

Procedure

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

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

Analyze

Use your data to complete the following steps. Record your answers in or below your data table.

1. Calculate the moles of NaCl contained in one level teaspoon.

$$\text{moles of NaCl} = \text{g NaCl} \times \frac{1\ \text{mol NaCl}}{58.5\ \text{g}}$$

2. Repeat Step 1 for the remaining compounds. Use the periodic table to calculate the molar mass of water and calcium carbonate.
3. Calculate 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.

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

$$\text{atoms of H} = \text{mol H} \times \frac{6.02 \times 10^{23}\ \text{atoms H}}{1\ \text{mol } H_2O}$$

Repeat for the other compounds in your table.

5. Which of the three teaspoon-sized samples contains the greatest number of moles?
6. Which of the three compounds contains the most atoms?

You're the Chemist!

The following small-scale activities allow you to develop your own procedures and analyze the results.

1. **Design It!** Can you count by measuring volume? Design and carry out an experiment to do it!
2. **Design It!** 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, $CaCO_3$.

Sample Data

	$H_2O(l)$	NaCl(s)	$CaCO_3(s)$
Mass (g)	4.30	5.09	9.68
Molar mass (g/mol)	18.0	58.5	100.1
Moles of compound	0.239	0.0870	0.0967
Moles of elements	0.239 O 0.478 H	0.0870 Na 0.0870 Cl	0.0967 Ca 0.0967 C 0.290 O
Atoms of elements	1.44×10^{23} O 2.88×10^{23} H	5.24×10^{22} O 5.24×10^{22} Cl	5.82×10^{22} Ca 5.82×10^{22} C 1.75×10^{22} O