

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.



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.

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

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

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

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

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.

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 \longrightarrow 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 .

$$\begin{aligned}\text{mass} &= 9.45 \text{ mol Al}_2\text{O}_3 \times \frac{102.0 \text{ g Al}_2\text{O}_3}{1 \text{ mol Al}_2\text{O}_3} \\ &= 964 \text{ g Al}_2\text{O}_3\end{aligned}$$

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 $\text{C}_{20}\text{H}_{42}$. 17. Calculate the mass, in grams, of 2.50 mol of iron(II) hydroxide.

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

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, $1 \text{ mol} = \text{molar mass}$, 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.

$$\text{moles of Na}_2\text{SO}_4 = 10.0 \text{ g} \times \frac{1 \text{ mol}}{142.1 \text{ g}} = 7.04 \times 10^{-2} \text{ mol}$$

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

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 : $1 \text{ mol} = 159.6 \text{ g } \text{Fe}_2\text{O}_3$
 Multiply the given mass by the conversion factor relating mass of Fe_2O_3 to moles of Fe_2O_3 .

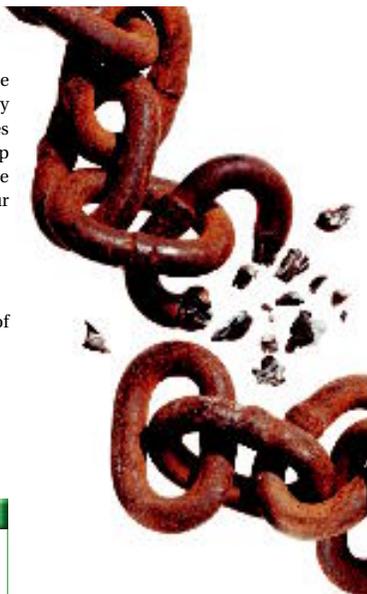
$$\begin{aligned} \text{moles} &= 92.2 \text{ g } \text{Fe}_2\text{O}_3 \times \frac{1 \text{ mol } \text{Fe}_2\text{O}_3}{159.6 \text{ g } \text{Fe}_2\text{O}_3} \\ &= 0.578 \text{ mol } \text{Fe}_2\text{O}_3 \end{aligned}$$

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.



Rust weakens an iron chain.

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.

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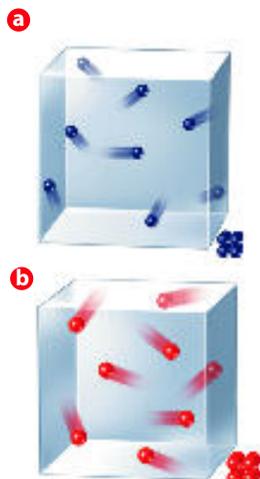


Figure 10.9 In each container, the volume occupied by the gas molecules is small compared with the container's volume, so the molecules are not tightly packed. **a** The molecules in this container are small. **b** This container can accommodate the same number of larger molecules.

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. 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?

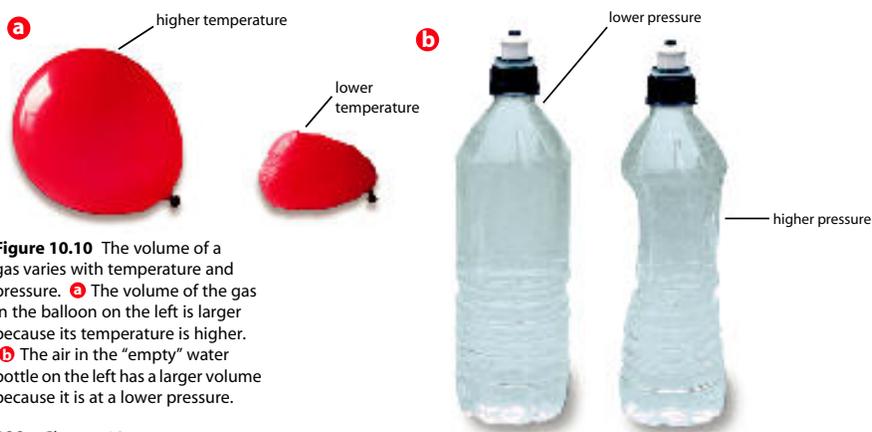


Figure 10.10 The volume of a gas varies with temperature and pressure. **a** The volume of the gas in the balloon on the left is larger because its temperature is higher. **b** The air in the “empty” water bottle on the left has a larger volume because it is at a lower pressure.

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 $22.4 \text{ L} = 1 \text{ mol}$ 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.

$$\text{volume of O}_2 = 0.375 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 8.40 \text{ L}$$



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

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
- $1 \text{ mol SO}_2 = 22.4 \text{ L SO}_2$

Unknown

- volume = ? L SO_2

Use the relationship $1 \text{ mol SO}_2 = 22.4 \text{ L SO}_2$ (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.

$$\text{volume} = 0.60 \text{ mol SO}_2 \times \frac{22.4 \text{ L SO}_2}{1 \text{ mol SO}_2} = 13 \text{ L SO}_2$$

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?
- $3.20 \times 10^{-3} \text{ mol CO}_2$
 - 3.70 mol N_2
21. At STP, what volume do these gases occupy?
- 1.25 mol He
 - $0.335 \text{ mol C}_2\text{H}_6$

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

$$\text{moles} = 0.200 \text{ L H}_2 \times \frac{1 \text{ mol H}_2}{22.4 \text{ L H}_2} = 8.93 \times 10^{-3} \text{ mol H}_2$$

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.

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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

- density = 1.964 g/L
- 1 mol (gas at STP) = 22.4 L

Unknown

- molar mass = ? g/mol

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.

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

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.

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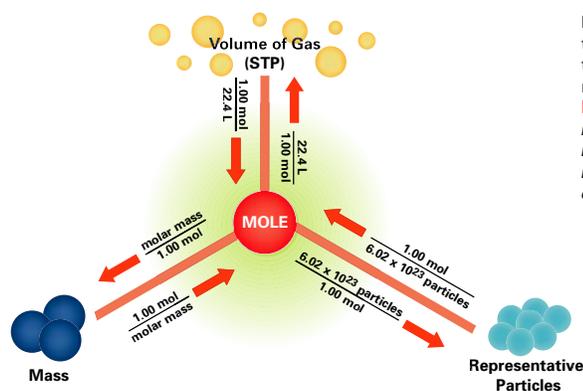


Figure 10.12 The map shows the conversion factors needed to convert among volume, mass, and number of particles. **Interpreting Diagrams** How many conversion factors are needed to convert from the mass of a gas to the volume of a gas at STP?

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

- Key Concept** Describe how to convert between the mass and the number of moles of a substance.
- Key Concept** What is the volume of one mole of any gas at STP?
- How many grams are in 5.66 mol of CaCO_3 ?
- Find the number of moles in 508 g of ethanol ($\text{C}_2\text{H}_6\text{O}$).
- Calculate the volume, in liters, of 1.50 mol Cl_2 at STP.
- The density of an elemental gas is 1.7824 g/L at STP. What is the molar mass of the element?
- 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).

- 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

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 (H₂O), and calcium carbonate (CaCO₃). Make a table similar to the one below.

	H ₂ O(l)	NaCl(s)	CaCO ₃ (s)
Mass (grams)			
Molar Mass (g/mol)			
Moles of each compound			
Moles of each element			
Atoms of each element			

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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 H₂O.

$$\text{moles of H} = \text{mol H}_2\text{O} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}}$$

Repeat for the other compounds in your table.

4. Calculate the number of atoms of each element present in the teaspoon-sized sample of H₂O.

$$\text{atoms of H} = \text{mol H} \times \frac{6.02 \times 10^{23} \text{ atoms H}}{1 \text{ mol H}_2\text{O}}$$

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, CaCO₃.