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 (O₂), 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, O₂ 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$_2$O$_3$). 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$_2$O$_3$
   
   Unknown
   - mass = ? g Al$_2$O$_3$

   The mass of the compound is calculated from the known number of moles of the compound. The desired conversion is moles \( \rightarrow \) mass.

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

   Determine the molar mass of Al$_2$O$_3$: 1 mol Al$_2$O$_3$ = 102.0 g Al$_2$O$_3$
   
   Multiply the given number of moles by the conversion factor relating moles of Al$_2$O$_3$ to grams of Al$_2$O$_3$.
   
   \[
   \text{mass} = 9.45 \text{ mol Al}_2\text{O}_3 \times \frac{102.0 \text{ g Al}_2\text{O}_3}{1 \text{ mol Al}_2\text{O}_3} \\
   = 964 \text{ g Al}_2\text{O}_3
   \]

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

   The number of moles of Al$_2$O$_3$ 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.

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**Practice Problems**

16. Find the mass, in grams, of $4.52 \times 10^{-3}$ mol C$_{20}$H$_{42}$

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

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Figure 10.8 These aluminum satellite dishes at the National Radio Astronomy Observatory near Soccoro, New Mexico are naturally protected from corrosion by the formation of a thin film of aluminum oxide (Al$_2$O$_3$).
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\(_2\)O\(_3\)). How many moles of iron(III) oxide are contained in 92.2 g of pure Fe\(_2\)O\(_3\)?

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

- **Known**
  - mass = 92.2 g Fe\(_2\)O\(_3\)
- **Unknown**
  - number of moles = ? mol Fe\(_2\)O\(_3\)

The unknown number of moles of the compound is calculated from a known mass of a compound. The conversion is mass \(\rightarrow\) moles.

**Calculate** Solve for the unknown.

Multiply the given mass by the conversion factor relating mass of Fe\(_2\)O\(_3\) to moles of Fe\(_2\)O\(_3\).

\[
\text{moles} = \frac{\text{mass (grams)}}{\text{molar mass (grams/mol)}} = \frac{92.2 \text{ g Fe}_2\text{O}_3}{159.6 \text{ g Fe}_2\text{O}_3} = 0.578 \text{ mol Fe}_2\text{O}_3
\]

**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\(_2\)O\(_3\) (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 \(\times\) \(10^{-1}\) g of boron.

19. Calculate the number of moles in 75.0 g of dinitrogen trioxide.
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 $6.02 	imes 10^{23}$ 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.

**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. The molecules in this container are small. This container can accommodate the same number of larger molecules.

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

**Check Point** What is meant by standard temperature and 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 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. 

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

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. 

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

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? 
   a. \(3.20 \times 10^{-4} \) 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\(_2\)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 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.

\[ \text{moles} = \frac{0.200 \text{ L H}_2 \times 1 \text{ mol H}_2}{22.4 \text{ L H}_2} = 8.93 \times 10^{-4} \text{ mol H}_2 \]
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.

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

\[
\text{grams mol}^{-1} = \frac{\text{grams}}{\text{L}} \times \frac{22.4 \text{ L}}{1 \text{ mol}}
\]

**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:

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

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?
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.

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 CaCO₃?

27. Find the number of moles in 508 g of ethanol (C₂H₅OH).

28. Calculate the volume, in liters, of 1.50 mol Cl₂ 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 (NH₃), sulfur dioxide (SO₂), chlorine (Cl₂), nitrogen (N₂), or methane (CH₄).

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.

Density
In Chapter 3 you learned that the densities of solids and liquids are measured in g/cm³ 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.

Assessment 10.2 Test yourself on the concepts in Section 10.2.

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### 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
- 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.

<table>
<thead>
<tr>
<th>Mass (grams)</th>
<th>Molar Mass (g/mol)</th>
<th>Moles of each compound</th>
<th>Moles of each element</th>
<th>Atoms of each element</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O(l)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaCl(s)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CaCO₃(s)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

1. Calculate the number of moles of NaCl contained in one level teaspoon.
   \[ \text{moles of NaCl} = \frac{\text{g NaCl}}{58.5 \text{ g/mol}} \times 1 \text{ mol NaCl} \]
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.
   \[ \text{moles of H} = \frac{\text{mol H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \times 2 \text{ mol H} \]
   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.
   \[ \text{atoms of H} = \frac{\text{mol H}}{1 \text{ mol H}_2\text{O}} \times 6.02 \times 10^{23} \text{ atoms H} \]
   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.