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 the question is asking for the mass of a mole of oxygen atoms (O), then the molar mass is 16.0 g. (Answers might include using the mass of one jelly bean and the total mass of all the jelly beans.)

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 headings, the boldfaced terms, and the visuals. List three things you expect to learn. After reading, state what you learned about each term you listed.

Objective
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 gas at STP.

Guide for Reading
Build Vocabulary

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

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
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.

Chemical Quantities
Section 10.2 (continued)

Sample Problem 10.5

Answers
16. \(4.52 \times 10^{-3}\) 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
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)

Relate
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
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.

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 (Al₂O₃). 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 Al₂O₃
   Unknown
   • mass = ? g Al₂O₃
   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 Al₂O₃: 1 mol Al₂O₃ = 102.0 g Al₂O₃
   Multiply the given number of moles by the conversion factor relating moles of Al₂O₃ to grams of Al₂O₃.
   \[
   \text{mass} = 9.45 \text{ mol Al}_2\text{O}_3 \times \frac{102.0 \text{ g Al}_2\text{O}_3}{1 \text{ mol Al}_2\text{O}_3} = 964 \text{ g Al}_2\text{O}_3
   \]

3. Evaluate Does the result make sense?
   The number of moles of Al₂O₃ 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 \(4.52 \times 10^{-3}\) mol \(C_{20}H_{42}\),
17. Calculate the mass, in grams, of 2.50 mol of iron(II) hydroxide.
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 (Na₂SO₄). 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 Na₂SO₄ into moles.

\[ \text{moles} = \frac{\text{mass (grams)}}{\text{molar mass (grams/mol)}} \]

The molar mass of Na₂SO₄ is 142.1 g/mol, so the number of moles of Na₂SO₄ is calculated this way:

\[ \text{moles of Na₂SO₄} = \frac{10.0 \text{ g}}{142.1 \text{ g/mol}} = 0.070 \text{ mol} \]

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

### Practice Problems

1. Calculate the number of moles in 1.00 × 10² 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.
The Mole–Volume Relationship

Discuss

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.

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 \(\frac{6.02 \times 10^{23}}{}\) 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?

Differentiated Instruction

Gifted and Talented

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)
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.

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

**SAMPLE PROBLEM 10.7**

**Calculating the Volume of a Gas at STP**

Sulfur dioxide ($\text{SO}_2$) 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 $\text{SO}_2$ gas at STP.

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

<table>
<thead>
<tr>
<th>Knowns</th>
<th>Unknown</th>
</tr>
</thead>
<tbody>
<tr>
<td>moles = 0.60 mol $\text{SO}_2$</td>
<td>volume = ? L $\text{SO}_2$</td>
</tr>
<tr>
<td>1 mol $\text{SO}_2$ = 22.4 L $\text{SO}_2$</td>
<td>Use the relationship 1 mol $\text{SO}_2$ = 22.4 L $\text{SO}_2$ (at STP) to write the conversion factor needed to convert moles to volume.</td>
</tr>
<tr>
<td>The conversion factor is $\frac{22.4 \text{ L} \text{SO}_2}{1 \text{ mol} \text{SO}_2}$.</td>
<td></td>
</tr>
</tbody>
</table>

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

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

**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. $3.20 \times 10^{-3}$ mol $\text{CO}_2$
   b. $3.70$ mol $\text{N}_2$

**21.** At STP, what volume do these gases occupy?
   a. $1.25$ mol $\text{He}$
   b. $0.335$ mol $\text{C}_2\text{H}_6$

The opposite conversion, from the volume of a gas at STP to the number of moles of gas, uses the same relationship: 22.4 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.

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

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**Differentiated Instruction**

**Gifted and Talented**

Have students design a way to measure the volume of $\text{CO}_2$ produced in the Teacher Demo on this page. Methods might include measuring how much water the filled bag displaces.

**Sample Problem 10.7**

**Answers**

20. a. $3.20 \times 10^{-3}$ mol $\text{CO}_2$ = 22.4 L $\text{CO}_2$/1 mol $\text{CO}_2$ = 7.17 $\times 10^{-2}$ L $\text{CO}_2$
   b. $3.70$ mol $\text{N}_2$ = 22.4 L $\text{N}_2$/1 mol $\text{N}_2$ = 82.9 L $\text{N}_2$

21. a. $1.25$ mol $\text{He}$ = 22.4 L $\text{He}$/1 mol $\text{He}$ = 28.0 L $\text{He}$
   b. $0.335$ mol $\text{C}_2\text{H}_6$ = 22.4 L $\text{C}_2\text{H}_6$/1 mol $\text{C}_2\text{H}_6$ = 7.50 L $\text{C}_2\text{H}_6$

**Practice Problems Plus**

**At STP, what volume is occupied by each of the following gases?**

a. $1.34$ mol $\text{SO}_2$ (30.0 L $\text{SO}_2$)
   b. $2.45 \times 10^{-3}$ mol $\text{H}_2\text{S}$ (5.49 $\times 10^{-2}$ L $\text{H}_2\text{S}$)
   c. $6.7$ mol $\text{H}_2$ (1.5 $\times 10^2$ L $\text{H}_2$)

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**Math Handbook**

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

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**TEACHER Demo**

**Molar Volume**

**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 $\text{CO}_2$) 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.

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**Answers to...**

- **Checkpoint**
  - $0^\circ$C and 101.3 KPa or 1 atmosphere (atm)

**Chemical Quantities** 301
Discuss

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

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

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.

\[
\text{molar mass} = \frac{\text{density at STP} \times \text{molar volume at STP}}{22.4 \text{ L/mol}}
\]

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 22.4 L/1 mol.

2. Calculate Solve for the unknown.

   \[
   \text{molar mass} = \frac{\text{density} \times \text{molar volume at STP}}{22.4 \text{ L/mol}}
   \]

   \[
   \text{molar mass} = \frac{1.964 \text{ g/L} \times 22.4 \text{ L}}{22.4 \text{ L/mol}}
   \]

   \[
   \text{molar mass} = 44.0 \text{ g/mol}
   \]

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?

Differentiated Instruction

Special Needs

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.
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).

### 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 CaCO₃

27. 13.0 mol C₂H₆O

28. 33.6 L Cl₂

29. 39.9 g/mol

30. gas A: 28.0 g, nitrogen

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.

### The Mole Road Map

Use Visuals

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.

### ASSESS

Evaluate Understanding

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

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.

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

Answers to...

Figure 10.12

The density of a gas is measured in grams per liter (g/L).
Counting by Measuring Mass

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 = 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 liquid and solid samples so that the container does not interfere with the measurement.
3. Design an experiment to do it!

For Enrichment
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.

## Materials
- chemicals shown in the table
- plastic spoon
- weighing paper
- watch glass 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>Compound</th>
<th>Mass (g)</th>
<th>Molar Mass (g/mol)</th>
<th>Moles of compound</th>
<th>Moles of each element</th>
<th>Atoms of each element</th>
</tr>
</thead>
<tbody>
<tr>
<td>H2O</td>
<td>4.30</td>
<td>18.0</td>
<td>0.239</td>
<td>0.0870 Na</td>
<td>0.0967 Ca</td>
</tr>
<tr>
<td>NaCl</td>
<td>5.09</td>
<td>58.5</td>
<td>0.0870 Na</td>
<td>0.0870 ClO</td>
<td>0.0967 Ca</td>
</tr>
<tr>
<td>CaCO3</td>
<td>9.68</td>
<td>100.1</td>
<td>0.0560 CaO</td>
<td>0.0675 ClO</td>
<td>1.75 × 10^23 O</td>
</tr>
</tbody>
</table>

### Sample Data

<table>
<thead>
<tr>
<th>Mass (g)</th>
<th>Moles of compound</th>
<th>Moles of elements</th>
<th>Atoms of elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>H2O</td>
<td>4.30</td>
<td>0.239</td>
<td>1.44 × 10^24 H</td>
</tr>
<tr>
<td>NaCl</td>
<td>5.09</td>
<td>0.0870 Na</td>
<td>5.24 × 10^24 O</td>
</tr>
<tr>
<td>CaCO3</td>
<td>9.68</td>
<td>0.0967 Ca</td>
<td>5.82 × 10^24 Ca</td>
</tr>
</tbody>
</table>

### Sample Calculations

1. Calculate the moles of NaCl contained in one level teaspoon.
   - moles of NaCl = g NaCl / 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.
   - moles of H = mol H2O × 2 mol H / 1 mol H2O
   - 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.
   - atoms of H = mol H × 6.02 × 10^23 atoms H / 1 mol 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?
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.