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Determination of the Molar Mass of Gases and Volatile Liquids

AP Chemistry Laboratory #3

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Introduction

The molar masses of compounds are used daily in the chemistry profession. The molar mass is defined as the mass, in grams, of 1 mole of any element or compound. How is molar mass determined and how is the molar mass of an unknown found? In this experiment, the molar masses of sample gases are determined directly and the molar masses of several volatile liquids will be calculated based on measurements of their vapor density.

Concepts

- Molar mass
- Ideal gas law
- Buoyancy

Background

The ideal gas law relates the four measurable properties of a gas (P, V, n, T). In this experiment, the ideal gas law will be used to determine the molar mass of gases and volatile liquids.

$$PV = nRT \qquad \textit{Equation 1}$$

The number of moles (n) of any pure substance is equal to its mass divided by its molar mass.

$$n = \text{mass/molar mass} \qquad \textit{Equation 2}$$

Substituting for n in Equation 1 and then rearranging produces the equation for the molar mass of a gas.

$$\text{molar mass (g/mol)} = \frac{\text{mass(g)} \times RT\text{(K)}}{P\text{(atm)} \times V\text{(L)}} \qquad \textit{Equation 3}$$

In Part 1, the mass of several "unknown" gases (X) is measured and compared to the mass of the same volume of oxygen. If two identical volumes of different gases are at the same temperature (T) and pressure (P), then their mass ratio must be equal to their molar mass ratio (Equation 4).

$$\frac{\text{mass of X}}{\text{mass of O}_2} = \frac{\text{molar mass of X}}{\text{molar mass of O}_2} \qquad \textit{Equation 4}$$

Rearranging:

$$\text{molar mass of X} = \frac{\text{mass of X}}{\text{mass of O}_2} \bullet \text{molar mass of O}_2 \qquad \textit{Equation 5}$$

When measuring the mass of a gas, the effect of buoyancy must be taken into account. Air, like water, exerts a positive or upward buoyant force on all objects. This force is compensated for in balances when massing liquids and solids. When massing gases, however, this force is *not* compensated for and is real. The apparent mass of a gas is less than the true mass of the gas.

$$\text{True mass of gas} = \text{apparent mass of gas} + \text{mass of air displaced} \qquad \textit{Equation 6}$$

In Part 1, the mass of each gas will be measured in a 60-mL gas syringe that has been evacuated of air and set to a fixed volume. The syringe is first massed with no gas in the fixed volume, then with a sample gas in the fixed volume. The effect of buoyancy is thus eliminated and the true mass of each gas, including air, can be determined directly.

In Part 2, the molar masses of several volatile liquids with boiling points well below the boiling point of water are determined. A small sample of the liquid is placed in a tared 15-mL plastic pipet and the pipet is then heated in boiling water to vaporize the liquid. The air and excess vapor escape, leaving the pipet filled only with the volatile liquid vapor at atmospheric pressure and at the temperature of boiling water. The pipet is then removed and cooled to condense the vapor.

Once cooled, the pipet is weighed. By massing the same pipet filled with deionized water, the volume of the pipet is calculated. The molar mass of the volatile liquid is then determined from Equation 3 using the mass of the condensed vapor, the volume of the pipet, the atmospheric pressure, and the temperature of the boiling water.

Experiment Overview

The purpose of this experiment is to determine the molar masses of various gases and volatile liquids. In Part 1, the gases are massed with a special gas syringe and their molar masses are determined by comparisons to data from oxygen measurements. In Part 2, liquids are volatilized and condensed in a fixed volume. The condensed vapor is massed and the liquid's molar mass is calculated from the experimental data.

Pre-Lab Questions

1. The following data represent a determination of the "molar mass" of a sample of air by comparison with the mass of oxygen as the reference gas. Assuming the air is 79% nitrogen, 20% oxygen, and 1% argon, and that these gases act as ideal gases, calculate both the experimental and theoretical "molar mass" of air. See Equation 4 in the *Background* section.

   Mass of evacuated syringe 40.687 g

	Air	O_2
Mass of syringe and gas	40.741 g	40.747 g
Mass of gas	0.054 g	0.060
Mass of gas/Mass of oxygen		1.00
Experimental molar mass		—
Theoretical molar mass		32.00

(*Hint:* For determining the theoretical molar mass of air, assume the percentages represent the mole fraction of each gas in the solution of air.)

2. A determination of the molar mass of methyl alcohol (CH_3OH) yielded the following data.

Temperature of boiling water bath	99.5 °C
Barometric pressure	738 mm Hg
Temperature of room temp. water bath	24.0 °C
Density of water at room temp.	0.9973 g/mL

	Trial 1
Mass of empty pipet	1.557 g
Mass of pipet and condensed methyl alcohol	1.571 g
Mass of pipet and water	16.001 g
Mass of condensed methyl alcohol	
Mass of water in filled pipet	
Volume of pipet	
Molar mass of methyl alcohol (experimental)	
Molar mass of methyl alcohol (theoretical)	

Using the data, fill in the rest of table. Calculate the molar mass of methyl alcohol using Equation 3 and compare this value to the actual molar mass of methyl alcohol. The volume of the pipet is equal to the volume of water in the pipet. Use the relationship of mass and density to determine this volume. Once the volume of the pipet is determined, equation 3 in the *Background* section can be used to calculate the molar mass of methyl alcohol.

Materials

Part 1

Luer-lock syringe, 60-mL, prepared with nail through plunger

Syringe tip cap, Luer-lock type

Gas cylinders, various
- Methane, CH_4
- Carbon dioxide, CO_2
- Oxygen, O_2

Balance, milligram (0.001-g precision)

Gas delivery bag with latex tubing and pinch clamp

Part 2

Acetone, CH_3COCH_3, 2 mL

Ethyl alcohol, CH_3CH_2OH, 2 mL

Isopropyl alcohol, $(CH_3)_2CHOH$, 2 mL

Beakers, 400-mL, 2

Boiling stones

Hot plate

Plastic tubing pipet holder

Test tube clamp

Beral-type pipets, super jumbo, narrow stem, 15-mL, 4

Balance, milligram (0.001-g precision)

Barometer

Pliers

Permanent marker

Thermometer

Ring stand

Scissors

Safety Precautions

Acetone, ethyl alcohol, and isopropyl alcohol are all flammable liquids and fire risks. Acetone and isopropyl alcohol are slightly toxic by ingestion and inhalation. Ethyl alcohol is made poisonous by the addition of denaturant—it cannot be made non-poisonous. If ammonia or chlorine gases are used in Part 1, release these gases in an efficient working fume hood. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Exercise care when working with the hot water bath. Wash hands thoroughly with soap and water before leaving the laboratory.

Procedure

Part 1. Molar Mass of Gas Samples

1. Push the plunger of the 60-mL syringe to the bottom of the specially prepared syringe and attach the syringe tip cap to the tip of the syringe.

2. Evacuate the syringe barrel: Pull the plunger to the 60-mL mark and place the nail in the prepared hole in the plunger so that the syringe plunger, when let go, returns to about the 50-mL mark. *Note:* This step requires two people—one person pulls the plunger out past the 50-mL mark and the other person then inserts the nail in the prepared hole.

3. Find the mass of the complete evacuated syringe assembly from step 2 to the nearest 0.001 g. Record the mass in the Part 1 Data Table.

4. Remove the syringe tip cap to allow air to enter the syringe. Replace the syringe tip cap and measure the mass of the complete syringe assembly filled with air. Record the mass in the Part 1 Data Table.

5. Remove the nail and the syringe tip cap. Depress the plunger to expel the air from the syringe.

6. Go to the gas delivery bag of oxygen in the fume hood and attach the syringe to the latex tubing. (This can be done by angling the tip of the syringe at 45 degrees to the end of the tubing, then working the tubing over the tip of the syringe.)

7. Release the pinch clamp on the gas delivery bag and draw oxygen into the syringe until the plunger is slightly past the 50-mL mark.

8. Insert the nail into the hole in the plunger, then push the plunger forward so the nail rests on the syringe barrel.

9. Reattach the pinch clamp to the latex tubing.

10. Hold the plunger in while releasing the syringe from the latex tubing. Immediately attach the syringe tip cap.

11. Measure the mass of the complete syringe assembly filled with oxygen and record the mass in the Part 1 Data Table.

12. Remove the syringe tip cap and expel the oxygen (use a fume hood when releasing gases that are unknown or poisonous).

13. Repeat steps 5–12 with the other assigned gases.

Part 2. Molar Masses of Volatile Liquids

1. Place a 400-mL beaker on the hot plate and add about 300 mL of water to the beaker, along with several boiling stones. Turn on the hot plate to boil the water.

2. Obtain three 15-mL jumbo Beral-type pipets. With pliers, pull the thin stems of each so that a very fine "capillary" tip is formed where the stem has been pulled (Figure 1).

3. Cut the pipet as shown in Figure 1 so that the capillary tip is less than 1-cm long.

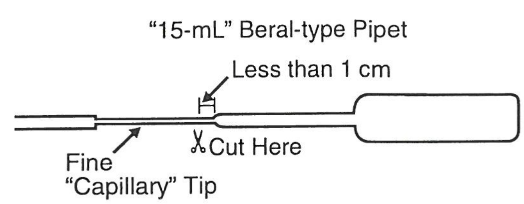

Figure 1.

4. Label the pipets #1, #2, and #3 with a permanent marker.

5. Mass each pipet to the nearest 0.001 g and record this mass in the Part 2 Data Table.

6. Draw 2–3 mL of the ethyl alcohol from the labeled bottle in the hood into each of the previously prepared and labeled pipets.

7. Insert the tips of the pipets containing the ethyl alcohol into the short piece of plastic tubing, then secure the tubing with a test tube clamp (Figure 2).

8. Lower the pipets into the boiling water bath. Make sure the entire bulb of each pipet is below the water line.

9. Heat for at least five minutes.

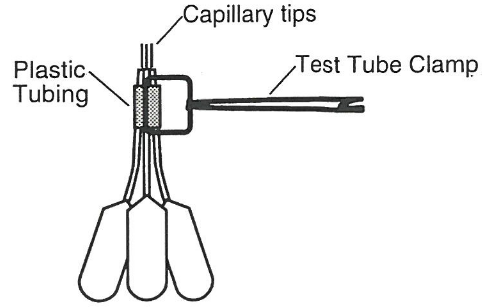

Figure 2.

10. Carefully remove the pipets from the water. Inspect each pipet. If any liquid remains in a pipet bulb, heat the entire assembly for another minute.

11. Cool the pipets by lowering the pipet assembly into a bath of room temperature water in a 400-mL beaker.
12. Record the temperature of the boiling water bath and the barometric pressure of the room in the Part 2 Data Table.
13. Dry the pipets with paper towels and mass each pipet, which now contains only the consensed vapor, to the nearest 0.001 g. Record these values in the Part 2 Data Table.
14. Fill a 250-mL beaker with room temperature deionized water.
15. Fill pipet #1 with the deionized water, then expel the water into the sink to flush the remaining ethyl alcohol from the pipet. Repeat this process several times.
16. To determine the volume of the pipet #1: Fill the pipet completely with deionized water, dry the outside, and mass the pipet and water. Record the mass in the Part 2 Data Table.
17. Repeat step 16 for pipets #2 and #3.
18. Repeat steps 2 through 13 for acetone and isopropyl alcohol.

Data Tables

Part 1

Mass of evacuated syringe ________ g

	Air	O_2*	Burner Gas	CO_2	Other
Mass of syringe and gas					
Mass of gas					
Mass of gas/Mass of oxygen					
Experimental molar mass					
Theoretical molar mass		32.0 g/mol			
Percent error					

*Oxygen is used as the reference gas for determining the molar mass of the other "unknown" gases.

Part 2

Temperature of boiling water bath __________ °C

Barometric pressure __________ mm Hg

Temperature of room temp. water bath __________ °C

Density of water at room temperature __________ g/mL

Jumbo Pipets

	Jumbo Pipet #1	Jumbo Pipet #2	Jumbo Pipet #3
Mass of empty pipet			
Mass of pipet and water			
Mass of water in filled pipet			
Volume of pipet			

Volatile Liquids

	Trial 1	Trial 2	Trial 3
Ethyl Alcohol			
Mass of pipet and condensed ethyl alcohol			
Mass of condensed ethyl alcohol			
Molar mass of ethyl alcohol			
Acetone			
Mass of pipet and condensed acetone			
Mass of condensed acetone			
Molar mass of acetone			
Isopropyl Alcohol			
Mass of pipet and condensed isopropyl alcohol			
Mass of condensed isopropyl alcohol			
Molar mass of isopropyl alcohol			

Post-Lab Questions and Calculations

(Use a separate sheet of paper to answer the following questions.)

Part 1

1. Why can the buoyancy force in this experiment be ignored?
2. Determine the mass of each gas in the syringe. Enter these values in the Part 1 Data Table.
3. How should the number of molecules trapped in the syringe compare between the various gases? Explain.
4. Determine the ratio of the mass of gas/mass of oxygen for each gas. Enter these values in the Part 1 Data Table.
5. How should the ratio of the mass of one molecule of gas/mass of one molecule of oxygen compare to the ratio of the mass of gas/mass of oxygen? Explain.
6. Use the molar mass of oxygen as a reference to determine the molar mass of each of the other gases tested in Part 1. Enter these values in the Part 1 Data Table.
7. Determine the accepted molar mass for each gas used (including the air value calculated in the *Pre-Lab* question #1).
8. Determine the percent error in your molar mass values.
9. How do the molar masses compare to the accepted values for each gas tested? Are there any patterns?
10. Which gases should have the greatest experimental uncertainty? Explain.

Part 2

1. Determine the mass of condensed, volatile vapor for each pipet trial and for each unknown in Part 2. Enter these values in the Part 2 Data Table.
2. Use the *CRC Handbook of Chemistry and Physics* to determine the density of water at the temperature of the room temperature water bath used in this experiment. Enter this density value in the Part 2 Data Table. Use this value and the mass of water in each filled pipet to calculate the volume of each pipet.
3. Determine the mass of the condensed volatile liquid for each run. Enter these values in the Part 2 Data Table.
4. Calculate the molar mass of the liquid used in each run and the average of the three runs for each volatile liquid.
5. Volatile liquids with lower boiling points often give better results then those with higher boiling points. Suggest a reason for this.
6. What effect would vapor condensation in the neck of the jumbo pipets have on the reported molar mass? How large an error might this introduce?
7. Some liquids have enough attractions between molecules to form dimers. (Dimers are molecules formed from the combination of the identical molecules, $A + A \rightarrow A_2$.) What effect would this have on the experimental molar mass?