*Experiment 3:*

# The Molar Volume of Gases

## *Purpose*

The purpose of this experiment is to determine the molar volume of a gas.

### *Learning Objectives*

Quantitatively measure the volume of oxygen gas produced from the decomposition of hydrogen peroxide.

Calculate the molar volume of oxygen gas at standard temperature and pressure using the ideal gas law and Dalton's law of partial pressures.

## Principles

### Molar Volume of an Ideal Gas

The **molar volume** is defined as the volume of one mole of a substance, or the volume per mole of a substance, $V/n$. The molar volume of a gas varies with temperature and pressure as the gas expands and compresses. We can calculate the molar volume of a gas from its density and molar mass at a given temperature and pressure.

$$\frac{V}{n} = \frac{\text{molar mass (g/mol)}}{\text{density (g/L)}} = \text{molar volume (L/mol)}$$

**Table 3.1:** Molar Volumes of Gases at 0 °C and 1 atm

| Gas | Density (g/L) | Molar Mass (g/mol) | Molar Volume (L/mol) |
|---|---|---|---|
| CO | 1.250 | 28.010 | 22.406 |
| $N_2$ | 1.250 | 28.013 | 22.403 |
| He | 0.178 | 4.003 | 22.429 |
| $H_2$ | 0.0899 | 2.016 | 22.429 |
| $CO_2$ | 1.977 | 44.010 | 22.262 |
| HCl | 1.639 | 36.461 | 22.244 |
| $O_2$ | 1.429 | 31.999 | 22.392 |

Table 3.1 clearly shows that for a variety of gases, at standard temperature and pressure (STP), the molar volume is very near 22.4 L/mol. In other words, at 0 °C and 1 atm, one mole of a gas occupies approximately 22.4 L. Two moles of gas would occupy 44.8 L. This is known as ***Avogadro's Law***: For a gas at constant temperature and pressure the volume is directly proportional to the number of moles of gas, $n$.

$$V \propto n \qquad \text{when pressure and temperature are constant}$$

The relationship between molar volume and pressure and temperature are described in Boyle's law and Charles' law as follows.

***Boyle's Law***: The volume of a fixed quantity of gas at a fixed temperature is inversely proportional to pressure. As the volume increases the pressure decreases.

$$P \propto (1/V) \qquad \text{when temperature and the amount of gas are constant}$$

***Charles' Law***: For a fixed quantity of gas at constant pressure, the volume increases as the temperature increases. A plot of volume versus temperature is a straight line and the intercept of this plot, when $T$ is measured in °C, is −273.15 °C. This is why absolute zero of temperature is defined as 0 K = −273.15 °C.
$T$ (in Kelvin) = $T$ (in °C) – 273.15.

$$V \propto T \qquad \text{when pressure and the amount of gas are constant}$$

Combining the gas laws of Boyle, Charles, and Avogadro, the ideal gas law is derived. The general equation given for the ideal gas law is an empirical law which holds approximately for all gases near atmospheric pressure and becomes increasingly accurate at low pressures.

$$PV = nRT \qquad \text{Ideal Gas Law}$$

The temperature, $T$, is the absolute temperature in units of Kelvin. $R$ is the universal gas constant that has the same value for all gases. $R$ can be estimated from the observation that one mole of any gas at 0 °C and 1 atm occupies approximately 22.4 L.

$$R = \frac{PV}{nT} = \frac{(1.0\ \text{atm})(22.4\ \text{L})}{(1\ \text{mol})(273\ \text{K})} = 0.08206\ \frac{\text{L atm}}{\text{mol K}}$$

The numerical value of $R$ depends on the units chosen for pressure and volume. Commonly used values of the gas constant, obtained through simple unit conversions, include the following:

$R = 0.08206\ \text{L atm mol}^{-1}\text{K}^{-1}$

$R = 1.987\ \text{cal mol}^{-1}\text{K}^{-1}$

$R = 8.314\ \text{J mol}^{-1}\text{K}^{-1}$

$R = 8.314\ \text{kg m}^2\ \text{sec}^{-2}\ \text{mol}^{-1}\text{K}^{-1}$

In this experiment you will determine the molar volume of oxygen at standard temperature and pressure (STP; 0 °C and 1 atm) and compare this value to the molar volume of 22.414 L/mol for an ideal gas at STP. To determine the molar volume of oxygen, the decomposition reaction of hydrogen peroxide, $H_2O_2$, to produce water and oxygen will be used.

$$2\,H_2O_2\,(aq) \xrightarrow{\text{catalyst}} 2\,H_2O\,(l) + O_2\,(g)$$

This reaction is slow and a catalyst is needed to speed up the reaction. The catalyst used in this experiment is $FeCl_3$; it speeds up the reaction without being consumed in the reaction itself. Because mass is always conserved in a reaction, the difference between the weight of reactants before the reaction takes place and the weight of products after the reaction goes to completion equals the mass of oxygen gas evolved. By measuring the volume of $O_2$ produced at a known temperature and pressure, we can calculate the molar volume of $O_2$ in L/mol (the volume of $O_2$ divided by the moles of $O_2$ produced).

The volume of oxygen gas produced in this reaction corresponds to the volume of water displaced by the gas. We must take into account that when any gas in a closed container is collected over liquid water, or is exposed to water, the water contributes to the total vapor pressure. Water evaporates until a saturated vapor results, that is, until opposing rates of evaporation and condensation of water molecules at the liquid surface reach a balance. According to Dalton's law of partial pressures, each gas exerts its own pressure regardless of the presence of other gases. In the collection of oxygen gas the total pressure is the sum of the partial pressure of oxygen and water.

$$P_{\text{TOTAL}} = P_{O_2} + P_{H_2O}$$

Thus, to determine the pressure of oxygen gas, a correction for the vapor pressure of water ($P_{H_2O}$) must be made.

$$P_{O_2} = P_{\text{TOTAL}} - P_{H_2O}$$

The state of a gas is defined by its pressure, volume, temperature, and the number of moles. For the change in state from $P_1$, $V_1$, $T_1$, $n_1$ to $P_2$, $V_2$, $T_2$, $n_2$ the following equations hold:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \text{at constant } n$$

This equation can be rearranged to: $V_2 = V_1\left(\frac{P_1}{P_2}\right)\left(\frac{T_2}{T_1}\right)$ at constant $n$

Thus, if we measure the volume of a gas to be $V_1$, at temperature $T_1$, and pressure $P_1$, the volume of the gas ($V_2$) can be calculated at any other temperature ($T_2$) and pressure ($P_2$). In this experiment, you will measure the volume of oxygen and hydrogen at the experimental temperature and pressure and you will then calculate the volume of each gas at STP (0 °C and 1 atm).

## Procedure

### Equipment

- Extra small test tubes, 12 × 75 mm, are supplied for the $FeCl_3$ catalyst. These test tubes must be returned after class.

### Chemicals

- Deionized water
- Hydrogen peroxide solution, 3%
- Iron(III) chloride solution, 3 M

### SAFETY

WEAR SAFETY GLASSES

If you spill any chemicals on your skin, flush with water immediately.

1. Assemble the apparatus as shown in Figure 3.1. Clean the glassware with soap and water, use deionized water (DI water) as the final rinse. For flask B use a 500 mL Erlenmeyer flask or a 500 mL filter flask. You may have one in your drawer. If not, borrow one from the stock room. Remove the yellow side arm from the filter flask and put a cork #3 in the hole on the side of the flask. Corks are provided in a bin along with the rubber stoppers and hoses.

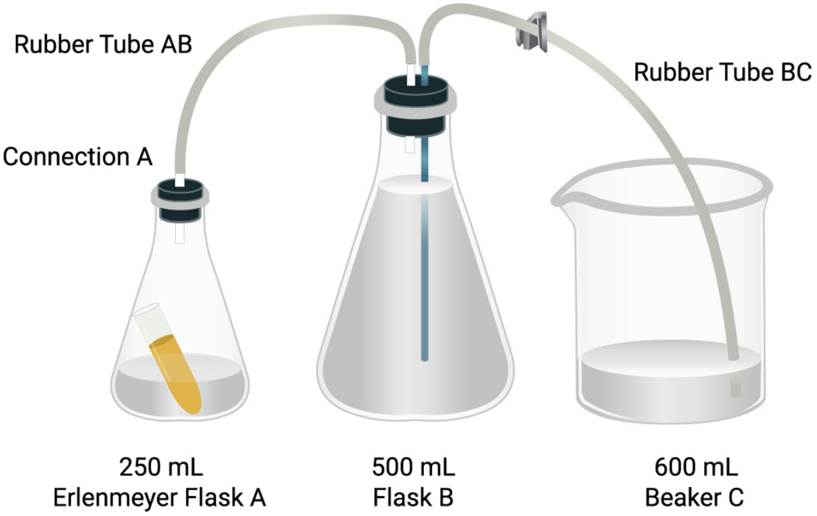

**Figure 3.1:** Note the position of the clamp—it is near the top of flask B.

2. Obtain an extra small test tube (12 × 75 mm) provided for this experiment. Test to make sure the small test tube remains standing at an angle in the 250 mL flask. If it drops completely horizontal inside the flask, obtain another flask or you can secure a metal paper clip to the open end of the test tube to prevent this.

3. Measure 20 mL of 3% $H_2O_2$ solution in a graduated cylinder and pour it into a 250 mL Erlenmeyer flask (flask A in the figure). Measure and pour approximately 4.5 mL of 3 M $FeCl_3$ into a 12 × 75 mm test tube (a very small test tube provided for this experiment). Holding flask A at an angle, carefully slide the test tube into the flask or lower the test tube with a pair of tongs.

4. Weigh flask A and its contents (before you weigh flask A, make sure it is dry on the outside). Record your data in Report Table 3.1 in the Report Sheet.

5. Fill the 500 mL Erlenmeyer flask B to the neck with tap water. Fill the 600 mL beaker C about one-third full with tap water. Disconnect the rubber stopper at point A and attach a pipette bulb. With the pinch clamp open, force air into rubber tube AB by squeezing the pipette bulb, forcing water from flask B into the beaker C. Raise and lower beaker C to move water back and forth through rubber tube BC to remove all air bubbles. With water halfway up the neck of flask B, reconnect the stopper at point A. Push down firmly on the stopper on flask A and on the stopper on flask B.

6. With the pinch clamp open, test the apparatus for leaks, as follows. Raise beaker C as high as possible without removing tubing from the beaker. The water level in flask B should move a little and then remain fixed. If the water level continues to change, a leak is present. Do not proceed until the leak is fixed.

7. Equalize the pressures inside and outside flask B by raising the beaker until the water level in the beaker and in flask B are the same. While one student holds the beaker to keep the water levels the same, another student closes the pinch clamp, positioned near the stopper on flask B. Pour out all the water in beaker C. Replace the tube in the beaker and open the pinch clamp. A little water will flow out and should be retained in the beaker.

8. Fill a trough half full with water.

9. Check to make sure the pinch clamp is open. Tip flask A carefully so that the $FeCl_3$ catalyst spills out of the test tube into the hydrogen peroxide solution (avoid getting the solution on the stopper). Place the flask in the water trough. Swirl the flask to mix the reactants. Note and record any changes that you see. As the oxygen is released, water is forced from flask B into beaker C. Feel the bottom of flask A. Is heat released or absorbed in the reaction? Note the level of the water in beaker C. Swirl the flask some more. When the reaction stops, and the water level remains constant, you can assume the decomposition of the hydrogen peroxide is complete.

10. Let the solution sit for 15 minutes to allow the temperature of the solutions in flasks A and B to equilibrate to room temperature.

11. Raise beaker C until the water levels in flask B and beaker C are equal, at which point another student closes the pinch clamp on tube BC. Measure the volume of oxygen produced by carefully pouring the water in beaker C into a graduated cylinder. Record the volume of water displaced by the oxygen gas.

12. To measure the temperature, use a digital thermometer. Remove the cap. Press the “on” button and check that the display reads °C rather than Fahrenheit. To measure the temperature, the °C display should not blink.

    > **NOTE:** When the °C display blinks, the thermometer is in hold position and will not record the temperature measured or any changes in temperature. If the °C display blinks, press the hold button to stop the blinking. When the °C display does not blink, the thermometer is set to measure the temperature.

13. To measure the temperature of the gas produced in the reaction, loosen the stopper on flask A and insert a thermometer into the solution. Measure and record the temperature of the solution in flask A. Similarly, measure and record the temperature of the solution in flask B. The average of these two temperatures represents the temperature of the oxygen gas.

14. Disconnect flask A at point A, dry the outside of flask A, weigh and record the flask and its contents.

15. Obtain the barometric pressure from your instructor. You can also obtain the value on the web: weather.com. For the local barometric pressure enter the name of the city or the zip code.

16. All waste is disposed of in the appropriately labeled waste container located in the hood.

17. Return the small test tubes provided for this experiment.

## Waste Disposal

Dispose of all waste in the appropriately labeled waste container in the hood.

Name: ______________________________

Section: __________ Date: __________

**Report Table 3.1:** Decomposition of $H_2O_2$ Using $FeCl_3$ as a Catalyst Measurements

| | |
|---|---|
| Mass of flask A and its contents before reaction (g) | ______________ |
| Mass of flask A and its contents after reaction (g) | ______________ |
| Temperature of the gas in flask A (°C) (which equals to the temperature of the solution in flask A) | ______________ |
| Temperature of the gas in flask B (°C) (which equals to the temperature of the solution in flask B) | ______________ |
| Average temperature of the gas (°C) and (K) | ______________ |
| Volume of oxygen gas collected (mL) | ______________ |
| Barometric pressure (atm) | ______________ |
| Water vapor pressure at the average temperature of the gas (atm) (see Appendix, interpolate the value)* | ______________ |

*See Appendix Table 3.3 at the end of this experiment. Show your calculation of the water vapor pressure, including interpolation, below Report Table 3.2.

## Calculations

Completed calculations must be shown to the instructor before you leave the lab.

**Report Table 3.2:** Molar Volume of Oxygen Calculations

| | |
|---|---|
| Mass of oxygen (g)* | ______________ |
| Moles of oxygen | ______________ |
| Average temperature of gas (absolute, K) | ______________ |
| Pressure of oxygen, $P_{O_2}$ (atm) | ______________ |
| Volume of $O_2$ at STP (calculate this from your data) (L) | ______________ |
| Molar volume of $O_2$ at STP (L/mol) | ______________ |
| % error** | ______________ |

*The mass of oxygen gas is very small. Keep 4 decimal places throughout calculations.
**Compare the molar volume of $O_2$ (g) at STP that you obtained to the actual value, 22.4 L/mol.

**CALCULATIONS:** Show your calculations, including UNITS throughout your calculations, for Average Temperature; Water vapor pressure at the average temperature; Mass of Oxygen gas produced; Moles of Oxygen gas, $O_2$ (g); Pressure of $O_2$ (g); Volume of $O_2$ (g) at STP; Molar Volume of $O_2$ (g) at STP (compare your value to the actual value, 22.4 L/mol).

## Review Questions

1. Explain the function of $FeCl_3$ in the decomposition of $H_2O_2$. If we would have used 5.0 mL of 3 M $FeCl_3$ rather than 4.5 mL of 3 M $FeCl_3$, would it have made any difference in the amount of oxygen gas collected? Explain why or why not.

2. In this experiment, some oxygen gas will dissolve in water. How will this affect the molar volume calculated? Would the molar volume be too high or too low?

3. In step 14, after the reaction has gone to completion, if the outside of flask A was wet when you weighed it, how would this affect your results? Would this cause the measured molar volume of oxygen gas to be higher or lower? Explain why.

4. If 0.65 g of a gas occupies 275 mL at 26 °C and 1.10 atm, calculate the molar mass of the gas.

5. The density of a gas is 0.980 g/L at 25 °C and 365 torr. Calculate the molar mass of the gas.

6. The density of dry air is 1.2929 g/L at STP. Calculate the average molar mass of air. Is this a reasonable value? Explain why or why not.

7. In this experiment, after the reaction went to completion, we let the solutions sit to equilibrate the temperature in flasks A and B. If you did not do this and the gas temperature was too high (in the reaction flask A) when you measured the volume of water displaced in the reaction, would the molar volume calculated be too high or too low? Explain why.

8. Under what conditions do real gases behave most like ideal gases? Does the air in a room at room temperature (22 °C) and pressure (1 atm) behave ideally?

9. The molar volume for $CO_2$ and HCl are 22.262 L/mol and 22.244 L/mol, respectively. Why are these molar volumes slightly less than the ideal molar volume, 22.4 L/mol?

10. Hydrogen gas can be produced from the reaction of calcium hydride and water.

$$\mathrm{CaH_2\,(s) + 2\,H_2O\,(l) \longrightarrow 2\,H_2\,(g) + Ca(OH)_2\,(aq)}$$

How many grams of calcium hydride are needed to produce 2.5 L of hydrogen gas, collected over water at 26 °C and 760. torr total pressure?

11. A series of measurements are made in order to determine the molar mass of an unknown gas. First, a large flask is evacuated and found to weigh 134.567 g. It is then filled with the gas to a pressure of 735 torr at 31 °C and reweighed; its mass is now 137.456 g. Finally, the flask is filled with water at 31 °C and found to weigh 1067.9 g. (The density of the water at this temperature is 0.997 g/mL.) Assuming that the ideal gas law equation applies, calculate the molar mass of the unknown gas.

12. A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed of 1.5 mol percent $CO_2$, 18.0 mol percent $O_2$, and 80.5 mol percent Ar.

    a. Calculate the partial pressure of $O_2$ in the mixture if the total pressure of the atmosphere is to be 745 torr.

    b. If this atmosphere is to be held in a 120 L space at 295 K, how many moles of $O_2$ are needed?

13. Many compounds used in the fragrance industry are derived from plant extracts. One step in identifying a desired compound is the determination of its molar mass. The volatile organic compound geraniol is a component of oil of roses. The mass density of the vapor at 260 °C and 103 Torr is 0.480 g/L. What is the molar mass of geraniol?

## Lab Report 3

For this experiment, the lab report should summarize the data and results in your own words. Write the discussion section in paragraph form and discuss the following:

Explain what was measured to determine the molar volume of $O_2$ (g). Include the balanced equation for the reaction after you introduce the reaction. Indicate the catalyst that was used and explain its purpose. Explain how you determined the volume of $O_2$ (g) at STP with the equation you used. Explain how you determined the molar volume of $O_2$ (g) at STP.

From your experimental data, state the molar volume of $O_2$ gas at STP that you determined (including units); state the accepted value at STP, 22.4 L/mol, and the percent error in the experimental molar volume relative to 22.4 L/mol.

**Sources of Error:** Review how the experiment was done and look at the data/results to state possible sources of error and how these errors affect the results:

1. State the experimental molar volume of $O_2$ (g) at STP and indicate if it is higher or lower than the accepted molar volume of $O_2$ (g), 22.4 L/mol, at STP.
2. State possible sources of error in determining the moles of $O_2$ (g) and explain if a given error caused your experimental molar volume at STP to be higher or lower than the accepted value at STP.

   When weighing the flask after the reaction, if the flask was not completely dry on the outside, explain how this would affect the mass, the moles and the molar volume of $O_2$ (g).
3. State possible sources of error in determining the volume of $O_2$ (g) and explain if a given error caused your experimental molar volume at STP to be higher or lower than the accepted value at STP.

   Explain the affect of oxygen dissolved in water on the volume of water displaced and on the experimental molar volume of $O_2$ (g).

   Explain how a difference in temperature affects the pressure of oxygen gas, the volume of water displaced and the experimental molar volume of $O_2$ (g).
4. Indicate which source of error is consistent with the experimental molar volume at STP relative to the known value.

For the conclusion, restate the objective and purpose of the experiment. Summarize major findings and how you would fix any of the errors that may have occurred.

The original data recorded in your laboratory notebook must be scanned and uploaded as part of your lab report.

## Preparation for Experiment 4

Before lab you MUST:

1. Read Experiment 4 procedure
2. Watch the two videos posted on LabFlow
3. Complete the PRE-LAB quiz
4. Complete the PRE-LAB Assignment

# Appendix

**Table 3.3:** Vapor Pressure of Water at Selected Temperatures

| Degrees Centigrade | Vapor Pressure (torr) | Degrees Centigrade | Vapor Pressure (torr) |
|---|---|---|---|
| –10 | 2.149 | 30 | 31.824 |
| –5 | 3.163 | 35 | 42.175 |
| 0 | 4.570 | 40 | 55.320 |
| 5 | 6.543 | 45 | 71.881 |
| 10 | 9.209 | 50 | 92.510 |
| 15 | 12.788 | 55 | 118.04 |
| 16 | 13.634 | 60 | 149.38 |
| 17 | 14.530 | 65 | 187.54 |
| 18 | 15.477 | 70 | 233.7 |
| 19 | 16.477 | 75 | 289.1 |
| 20 | 17.535 | 80 | 355.1 |
| 21 | 18.650 | 85 | 433.6 |
| 22 | 19.827 | 90 | 525.8 |
| 23 | 21.068 | 95 | 633.9 |
| 24 | 22.377 | 100 | 760.00 |
| 25 | 23.756 | 125 | 1740 |
| 26 | 25.209 | 150 | 3570 |
| 27 | 26.739 | 175 | 6694 |
| 28 | 28.349 | 200 | 11659 |
| 29 | 30.043 | 300 | 64432 |