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Calculating the Molar Volume of Carbon Dioxide

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Purpose of the Experiment

Determine the molar volume of carbon dioxide by measuring the mass and volume of a carbon dioxide gas sample at laboratory temperature and pressure.

Background Required

You should be familiar with basic laboratory techniques for measuring mass, volume, and temperature, and with the concepts associated with gas laws and gas stoichiometry.

Background Information

Molecules of an **ideal gas** are small and nonpolar; they have no attractive or repulsive forces between them. **Avogadro's law** states that, at comparable temperatures and pressures, equal volumes of ideal gases contain equal numbers of molecules. Thus, at the same temperature and pressure, one mole of any ideal gas occupies the same volume as one mole of any other ideal gas. This volume is called the **molar volume (V_m)** of an ideal gas. For convenience of reference, we identify **standard temperature and pressure (STP)** as a temperature of 273 kelvin (K) and a pressure of one atmosphere (P = 1 atm, which is also equal to 760 torr). At STP, the molar volume of any ideal gas is 22.4 L.

We can determine the molar volume of carbon dioxide (CO_2), which behaves almost ideally, by collecting a CO_2 sample under non-standard conditions and applying the combined gas law to determine its volume at STP. The **combined gas law** (Equation 1) indicates the relationship among pressure (P), volume (V), and temperature (T) of a gas sample under two sets of conditions, "1" and "2."

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \text{(Eq. 1)}$$

We can identify condition "1" as the volume, temperature, and pressure of the CO_2 sample at laboratory conditions and condition "2" as STP. We can then solve Equation 1 for V_2, which becomes V_{STP}, as shown in Equation 2, and calculate the CO_2 sample volume at STP.

$$V_{STP} = \frac{P_1V_1T_{STP}}{T_1P_{STP}} \qquad \text{(Eq. 2)}$$

 01 00 99 — 15 14 13 12 11 10 9 8 7 6 5 4 3 2 1

From V_{STP} and the mass of the collected gas, we can calculate the molar volume of the sample at STP.

Example

Problem You determine that in a laboratory where $T = 22.2$ °C and $P = 30.05$ in. Hg, the mass of a dry, empty stoppered flask is 82.151 g. The mass of the same stoppered flask filled with CO_2 is 82.237 g. When completely filled and tightly stoppered, the flask contains 165 mL of water. According to the *CRC Handbook of Chemistry and Physics*, the density of air at 22.2 °C and 30.05 in. Hg is 1.201 g/L. Use these data to calculate the molar volume of CO_2 at STP.

Solution *(1) Express the laboratory pressure (P_1) in atmospheres, temperature (T_1) in kelvins, and the sample volume (V_1) in liters, recalling that 25.4 mm Hg = 1 in. Hg; 1 atm = 760 torr = 760 mm Hg; T, K = T, °C + 273; and 1 L = 1000 mL.*

$$P_1\text{, atm} = 30.05\text{ in. Hg}\left(\frac{25.4\text{ mm Hg}}{1\text{ in. Hg}}\right)\left(\frac{1\text{ atm}}{760\text{ mm Hg}}\right) = 1.00\text{ atm}$$

$$T_1\text{, K} = 22.2\text{ °C} + 273 = 295\text{ K}$$

$$V_1\text{, L} = 165\text{ mL}\left(\frac{1\text{ L}}{1000\text{ mL}}\right) = 0.165\text{ L}$$

(2) Determine the equivalent volume of the collected CO_2 at STP (V_{STP}), using Equation 2.

$$V_{STP} = \frac{P_1V_1T_{STP}}{T_1P_{STP}} = \frac{(1.00\text{ atm})(0.165\text{ L})(273\text{ K})}{(295\text{ K})(1\text{ atm})} = 0.153\text{ L}$$

(3) Calculate the mass of air in the flask, m_a, using Equation 3.

$$m_a = (\text{volume of flask, L})(\text{density of air, g/L}) = (0.165\text{ L})(1.201\text{ g/L}) = 0.198\text{ g} \quad \text{(Eq. 3)}$$

(4) Calculate the mass of CO_2 in the flask.

The mass from the first weighing of the flask (m_1) is the mass of the stoppered flask (m_f) plus the mass of the air in the flask (m_a);

$$m_1 = m_f + m_a$$

The mass from the second weighing of the flask (m_2) is the mass of the flask (m_f) plus the mass of CO_2 gas contained in the flask (m_{CO_2});

$$m_2 = m_f + m_{CO_2}$$

Thus, the difference between the second and first masses is the mass of CO_2 minus the mass of air;

$$m_2 - m_1 = m_{CO_2} - m_a$$

So, we can calculate the mass of CO_2 in the flask using Equation 4.

$$m_{CO_2} = m_2 - m_1 + m_a = 82.237\text{ g} - 82.151\text{ g} + 0.198\text{ g} = 0.284\text{ g} \quad \text{(Eq. 4)}$$

(5) Calculate the number of moles of CO_2 in the sample, n, using Equation 5.

$$n = (m_{CO_2}\text{, g})\left(\frac{1\text{ mol }CO_2}{44.01\text{ g }CO_2}\right) = (0.284\text{ g }CO_2)\left(\frac{1\text{ mol }CO_2}{44.01\text{ g }CO_2}\right) = 6.45\times10^{-3}\text{ mol }CO_2 \quad \text{(Eq. 5)}$$

(6) Calculate the molar volume of CO_2 at STP, V_m, using Equation 6.

$$V_m = \frac{V_{STP}}{n} = \frac{0.153\ \text{L}}{6.45 \times 10^{-3}\ \text{mol}\ CO_2} = 23.7\ \text{L / mol}\ CO_2 \qquad \text{(Eq. 6)}$$

In This Experiment

You will collect a CO_2 sample by allowing dry ice (solid CO_2) to **sublime**, that is, to pass directly from the solid to the gas phase, in a flask under conditions where you can determine the temperature and pressure. The subliming CO_2 will force all the air from the flask because CO_2 is considerably more dense than air. From the mass and volume of the collected CO_2, you will calculate the molar volume of CO_2.

Procedure

Caution: *Wear departmentally approved safety goggles while doing this experiment. Always use caution in the laboratory. Many chemicals are potentially harmful. Prevent contact with your eyes, skin, and clothing. Avoid ingesting any of the reagents.*

Note:
- *If you are not familiar with the proper technique for using the balances available in your laboratory, ask your laboratory instructor for assistance.*
- *Record all data on your Data and Observations sheet.*
- *Complete duplicate or triplicate determinations of this experiment, as indicated by your laboratory instructor.*

1. Weigh a dry, stoppered 125-mL Erlenmeyer flask to the nearest 0.001 g. Record the mass (m_1) of the flask, stopper, and contained air, in the column headed "determination 1". Then, remove the stopper from the flask.

Caution: *Dry ice can cause frostbite. Do not handle dry ice with bare hands.*

Note: *In Step 2, do not insert the stopper into the flask as long as any solid CO_2 is present.*

2. Obtain a 3/4-in. cube of dry ice from your laboratory instructor. Using forceps, place the solid CO_2 in the unstoppered flask and allow the dry ice to sublime.

When the solid has ***completely*** sublimed, tightly stopper the flask.

3. Weigh the flask, stopper, and CO_2 gas. Record this mass (m_2) in the "determination 1" column.

4. To do a second determination, pour the CO_2 from your flask by slowly inverting it. When the CO_2 has been completely replaced by air, repeat Steps 1–3, recording your data in the column labeled "determination 2".

5. To do a third determination, repeat Step 4, recording your data in the column labeled "determination 3".

6. When you have completed the required number of determinations, unstopper the flask and fill it to the brim with tap water. Stopper the flask tightly, displacing some of the water. Then wipe any excess water from the outside of the flask.

7. Unstopper the flask. Carefully pour the water from the flask into a 100-mL graduated cylinder. Fill the calibrated portion of the cylinder and record this volume. Empty the cylinder into the sink. Pour the water remaining in the flask into the cylinder. Read the new volume. Add the two volumes together and report this sum as the total volume, to the nearest milliliter, of the flask. Record this volume as the volume of the flask (V_1).

8. Record the barometric pressure (P_1). Be sure to indicate the units used.

9. Record the laboratory temperature (T_1), to the nearest 0.1 °C.

10. Obtain from your laboratory instructor the density of air at the laboratory temperature and pressure. Record this air density.

Caution: *Wash your hands thoroughly with soap or detergent before leaving the laboratory.*

name partner section date

Post-Laboratory Questions

Use the spaces provided for the answers and additional paper if necessary.

1. Calculate the percent error for your determination relative to the theoretical molar volume of 22.4 L/mol, using Equation 7.

$$\text{percent error, } \% = \frac{\text{experimental } V_m\text{, L / mol} - 22.4\text{ L / mol}}{22.4\text{ L / mol}}(100\%) \qquad \text{(Eq. 7)}$$

2. The success of this experiment depends on the total replacement of the air in the Erlenmeyer flask with CO_2. Such replacement can only occur if CO_2 is more dense than air. Give the data that you measured in the laboratory that verify that CO_2 is more dense than air.

3. You measured the volume of the CO_2 by filling your flask with water and tightly inserting the stopper.

(a) Why was it necessary to tightly stopper the water-filled flask and then remove the stopper before you emptied the water into a graduated cylinder for measurement?

(b) Would your calculated molar volume have been higher or lower if you had forgotten to stopper the flask before measuring the volume of the flask? Briefly explain.

4. Your source of CO_2 in this experiment was dry ice. Would your experimentally determined V_m of CO_2 have been too high or too low if you had:

(a) used a piece of dry ice that was too small to produce enough CO_2 to completely fill the flask? Briefly explain.

(b) used a large piece of dry ice and not waited for the entire sample to sublime before you stoppered the flask and proceeded with the experiment? Briefly explain.

name	partner	section	date

Data and Observations

	determination 1	2	3
mass of flask, stopper, and air (m_1), g	______	______	______
mass of flask, stopper, and CO_2 (m_2), g	______	______	______
volume of flask (V_1), mL	______	______	______
laboratory barometric pressure (P_1) (indicate units)		______	
laboratory temperature (T_1), °C		______	
density of air at T_1 and P_1, g/L		______	

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Calculations and Conclusions

Show your calculations in the spaces provided. Remember to include units with all calculated results.

Do the following calculations. Where applicable, complete a second calculation for each determination you performed.

1. Convert your laboratory barometric pressure measurement, P_1, to atmospheres.

2. Convert the laboratory temperature, T_1, to kelvins.

3. Determine the volume, in liters, of your flask at laboratory temperature and pressure.

4. Determine the volume (in liters) of your CO_2 gas sample under standard conditions, V_{STP}, using Equation 2.

5. Calculate the mass of air contained in the flask (m_a), using Equation 3.

6. Calculate the mass of CO_2 contained in the flask (m_{CO_2}), using Equation 4.

7. Find the number of moles of CO_2 in your sample (n), using Equation 5.

8. Using Equation 6, calculate the experimentally determined molar volume of CO_2 (V_m) at STP.

9. If you did multiple determinations, calculate the average molar volume of CO_2. Add the experimentally determined molar volumes together and divide by the number of determinations.

name partner section date

Pre-Laboratory Assignment

1. What precaution should you take when you work with dry ice?

2. Briefly explain the meaning of the following sentences.

(a) Dry ice sublimes.

(b) CO_2 behaves almost ideally.

3. How are you going to determine the following in the laboratory:

(a) the volume of the CO_2 sample?

(b) the mass of the air in the flask?

4. A student performing this experiment collected the following data:

laboratory temperature = 23.0 °C;
laboratory pressure = 30.05 in. Hg;
mass of a dry, empty stoppered flask = 82.715 g;
mass of the same stoppered flask filled with CO_2 = 82.795 g;
volume of water contained by stoppered flask = 144 mL;
density of air under the laboratory conditions = 1.193 g/L

(a) Using the student's data, calculate the experimentally determined molar volume of CO_2.

(b) Calculate the percent error in the student's experimentally determined molar volume of CO_2.

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