THE MOLAR VOLUME OF CARBON DIOXIDE

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The Problem to be Investigated: The molar volume of carbon dioxide will be determined by measuring the mass of a sample of carbon dioxide vapor of known volume.

The Nature of This Investigation: An unknown mass of solid carbon dioxide is vaporized in a tarred flask. The flask containing carbon dioxide is weighed and, by approximate calculations, the mass of the carbon dioxide vapor is found. Correction to standard conditions (STP) permits calculation of the molar volume of carbon dioxide.

BACKGROUND INFORMATION

Avogadro's law states that equal volumes of ideal gases under identical conditions of temperature and pressure contain equal numbers of molecules. Thus, one mole of an ideal gas would occupy the same volume as one mole of any other ideal gas at the same temperature and pressure. The volume occupied by one mole of a gas is the molar volume of that gas under the given conditions of temperature and pressure. At standard conditions, one atmosphere pressure and 0°C, the molar volume of any ideal gas is 22.414 liters.

The ideal gas law in Equation (1) expresses the relation between pressure, *P*, volume, *V*, number of moles of gas, *n*; and the absolute temperature, *T*, of the gas.

$$PV = nRT \qquad \text{(Eq. 1)}$$

R is the gas law constant and has the value of 0.0821 liter-atm deg^{-1} mole^{-1}, for all gases independent of temperature, volume, pressure, or number of moles.

If a given number of moles of a gas are contained in a volume, V_1; at pressure, P_1; and temperature, T_1, the ideal gas law may be written

$$P_1V_1 = nRT \qquad \text{(Eq. 2)}$$

If the pressure, volume, and temperature of the gas are changed to P_2, V_2, and T_2, while keeping the number of moles of gas the same, the ideal gas law for the new set of conditions is

$$P_2V_2 = nRT_2 \qquad \text{(Eq. 3)}$$

Equations (2) and (3) may be rearranged to

$$\frac{P_1V_1}{T_1} = nR \qquad \text{(Eq. 4)}$$

$$\frac{P_2V_2}{T_2} = nR \qquad \text{(Eq. 5)}$$

Because P_1V_1/T_1 and P_2V_2/T_2 are equal to the same constant, they are equal to each other; thus,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \text{(Eq. 6}$$

Equation (6) is used in this experiment to convert the laboratory conditions, P_1, V_1, and T_1 to standard conditions. The temperature appearing in Equation (6) must be expressed in Kelvins, which is found by adding 273.15 to any Celsius temperature. The pressure may be expressed in Torr or atmospheres, and volume in milliliters or liters.

As an illustration of the application of these equations for the determination of a molar volume, consider the following data. The volume of 0.250 g of carbon dioxide gas at a pressure of 754 torr and a temperature of 27°C was found to be 145 mL. Insertion of the appropriate data into Equation (6) gives

$$\frac{(754\_torr)(145\_mL)}{(273.15+27)K} = \frac{(760mmHg)(V_2)}{273K}$$

Equation (7) is rearranged to

$$V_2 = \frac{(273K)(754torr)(145mL)}{(300K)(760torr)} = 131mL$$

Then V_2 is found to be 131 mL, which is the gas volume corrected to standard conditions. The molar volume of the gas at standard conditions is found by dividing the volume of the gas by the number of moles of gas. The number of moles, *n*, of carbon dioxide is

$$n = \frac{mass\_CO_2,\_g}{molecular\_mass\_CO_2,\_g\_mole^{-1}}$$

$$n = \frac{0.250\,g}{43.99\,g\_mol^{-1}} = 5.68 \times 10^{-3}\,mol$$

Thus the molar volume, V_m, of carbon dioxide is

$$V_m = \frac{V_2}{n} = \frac{131\,mL}{5.68 \times 10^{-3}\,mole} = 2.31 \times 10^{4}\,mL\_mol^{-1}$$

$$V_m = 23.1\_L\_mol^{-1}$$

The percent error as compared to 22.4 liter mol^{-1} is found from Equation (11).

$$Percent\_error = \frac{\left|(\text{experimental} - \text{theoretical})\right|}{\text{theoretical}} \times 100$$

For V_m as found above,

$$Percent\_error = \frac{23.1 - 22.4}{22.4} \times 100 = 3\%$$

Without regard for the sign of the error.

For this experiment the volume of carbon dioxide gas is determined by finding the mass of a known volume of the gas at the prevailing laboratory temperature and pressure. The mass is determined by first weighing a flask filled with air then reweighing the flask filled with carbon dioxide gas. The flask is filled with carbon dioxide gas by allowing a piece of solid carbon dioxide, Dry ice, to vaporize in the flask. The vaporization process forces the air from the flask leaving the flask filled with carbon dioxide gas at the laboratory temperature and pressure. The volume of gas is the same as the volume of the flask.

The mass found for the first weighing of the flask, m_1, includes the mass of the flask, m_f, plus the mass of air, m_a, contained in the flask. Thus,

$$m_1 = m_f + m_a \qquad \text{(Eq. 1}$$

The mass found for the second weighing of the flask filled with carbon dioxide gas, m_2, is the mas of the flask, m_f, plus the mass of carbon dioxide gas contained in the flask, m_{CO2}. The expansion of the carbon dioxide gas is assumed to have expelled all air from the flask so that no air mass is included in the second weighing. Thus,

$$m_2 = m_f + m_{CO_2} \qquad \text{(Eq}$$

The difference between the second and first weighings is then

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

Equation (14) may be rearranged to calculate the mass of carbon dioxide gas,

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

The mass of carbon dioxide is then not equal to the difference between the two weighings but must be corrected by the mass of air contained in the flask. M_a may be found by multiplying the density of air at the temperature and pressure of the weighings by the volume of air contained in the flask.

$$m_a = (density\_of\_air)(volume\_of\_flask)$$

Notes on Equipment:

General: Balance, ±0.001 g
Individual: Cylinder, graduated, 100-mL; Flask, Erlenmeyer, 125-mL; Stopper, rubber, solid No. 4

Notes on Reagents:

Dry Ice (3/4" pieces), one per student determination (1 cu ft could conceivably serve about 100 students, with three determinations each; this is a very rough estimate).

Reference Table: Density of dry air

	Pressure H, cm							
Temp., deg.	72.0	73.0	74.0	75.0	76.0	77.0	Proportional parts	
10	0.00118	0.001198	0.001215	0.001231	0.001247	0.001264	17 cm	
11	178	193	210	227	243	259	0.1	2
12	173	190	206	222	239	255	0.2	3
13	169	186	202	218	234	251	0.3	5
14	165	181	198	214	230	246	0.4	7
							0.5	8
							0.6	10
							0.7	12
15	0.001161	0.001177	0.001193	0.001210	0.001226	0.001242	0.8	14
16	157	173	189	205	221	238	0.9	15
17	153	169	185	201	217	233	16 cm	
18	149	165	181	197	213	229		
19	145	161	177	193	209	225	0.1	2
							0.2	3
							0.3	5
							0.4	6
20	0.001141	0.001157	0.001173	0.001189	0.001205	0.001221	0.5	8
21	137	153	169	185	201	216	0.6	10
22	134	149	165	181	197	212	0.7	11
23	130	145	161	177	193	208	0.8	13
24	126	142	157	173	189	204	0.9	14
							15cm	
25	00.1122	0.001138	0.001153	0.001169	0.001185	0.001200	0.1	1
26	118	134	149	165	181	196	0.2	3
27	115	130	146	161	177	192	0.3	4
28	111	126	142	157	173	188	0.4	6
29	107	123	138	153	169	184	0.5	7
							0.6	9
							0.7	10
							0.8	12
30	0.001104	0.001119	0.001134	0.001150	0.001165	0.001180	0.9	13

EXPERIMENTAL PROCEDURE

1. By using a balance sensitive to 0.001 grams, carefully weigh a stoppered 125-mL Erlenmeyer flask to three decimal places. Record the mass of the flask and stopper on the Data Sheet.

2. Obtain from the laboratory instructor a piece of solid carbon dioxide, Dry Ice, about 3/4 inch wide.

   CAUTION: *Do not handle the Dry Ice with bare hands because of the danger of frost bite.*

3. Place the solid carbon dioxide in the unstoppered flask.

4. When all the solid has just disappeared or sublimed, stopper the flask tightly.

5. Weigh the flask, stopper, and carbon dioxide vapor to three decimal places. Record the mass on the Data Sheet.

6. After the flask, stopper, and contents have been weighed, fill the flask to the brim with water.

7. Stopper the flask tightly.

8. By using a towel or absorbent paper tissue, wipe the excess water from the outside of the flask.

9. Carefully pour the water from the flask into a 100-mL graduated cylinder. Record on a Data Sheet the volume of water. This is the volume of the flask.

10. Record on the Data Sheet the barometric pressure and the temperature of the laboratory.

11. Do a second and third determination and if time is available a third determination.

CALCULATIONS

1. Correct the volume of carbon dioxide gas to standard conditions.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{(Eq. 6}$$

where P_1 and T_1 are the laboratory pressure and temperature. V_1 is the volume of the flask, P_2 is 760 tor, T_2 is 273K, and V_2 is the volume corrected to standard conditions.

2. Calculate the mass of air contained in the flask.

   a. By using the Reference Table, determine the density of air at the temperature and pressure of this experiment.

   b. The mass of air, m_a, in the flask is the density multiplied by the volume of the flask.

$$m_a = \text{density, g mL}^{-1} \text{ x volume of flask, mL}$$

3. Calculate the mas of carbon dioxide gas, m_{CO2}, contained in the flask.

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

where m_2 - m_1 is the difference in the two weighings and m_a is the mass of the air in the flask.

4. Find the number of moles of carbon dioxide gas, n.

$$n = \frac{mass\_CO_2,\_g}{molecular\_mass\_CO_2,\_g\_mole^{-1}}$$

5. Calculate V_m, the molar volume.

$$V_m = \frac{V_2}{n}$$

DATA SHEET

Molar Volume of CO_2

Determinations:	1	2	3
Room temperature (T_1)	___________	___________	___________
Atmospheric pressure (P_1)	___________	___________	___________
Volume of flask (V_1), mL	___________	___________	___________
Mass of flask + air, m_1, g	___________	___________	___________
Mass of flask + CO_2, m_2, g	___________	___________	___________
Density of air at T_1, P_1 (d)	___________	___________	___________
Mass of air, m_a, g	___________	___________	___________
Corrected vol. of CO_2 (V_2), mL	___________	___________	___________
Molecular weight of CO_2	___________	___________	___________
Mass of CO_2	___________	___________	___________
Number of moles of CO_2	___________	___________	___________
Molar volume of CO_2 (V_m)	___________	___________	___________
Average value of V_m		___________	
Percent error		___________	

QUESTIONS

1. A student obtained a mass of 0.210 g CO_2 during his experiment. The atmospheric pressure was 751 torr and the laboratory temperature was 24°C. The flask used in the experiment had a volume of 146 mL. Calculate the molecular weight of CO_2 from the data (V_m = 22.414 liters mole^{-1} at STP).

2. What was the percent error from the results of the experiment in Question 1 above?

3. If, in Question 1 above, the molecular weight of CO_2 was supplied instead, what would have been the value of molar volume?

4. What is the percent error from the results of Question 3?

PRE-LABORATORY ASSIGNMENT

Molar Volume of CO_2

1. A mole of CO_2 has a mass of _________ (units).

   A mole of O_2 has a mass of _________.

   A mole of N_2O_5 has a mass of _________.

   A mole of C_2H_6 has a mass of ________.

   These substances are all gases at ordinary temperatures.

   32 grams of oxygen occupy a volume of 22.4 liters at 0°C and 1 atm pressure; thus:

   _________(units) of N_2O_5 occupy a volume of ________;

   _________ of C_2H_6 occupy a volume of ________;

   and, __________ of CO_2 occupy a volume of ________ at 0°C and 1 atm pressure.

2. Why use grams as a weight unit?

3. Why say mass instead of weight, when we say molecular weight rather than molecular mass?

4. Is a mole a molecular mass?