Activity 15 - The Molar Mass of a Gas

Objective

The purpose of this experiment is to determine the number of grams per mole of a gas by measuring the pressure, volume, temperature, and mass of a sample.

Terms to Know

- *Molar Mass* - The number of grams per mole of a substance.
- *Ideal Gas Law* - The relation between pressure, volume, the number of moles, and temperature of a gas: $PV = nRT$
- *Vapor Pressure of Water* - The pressure exerted by water vapor as it saturates the environment above a sample of liquid water. Water vapor pressures depend upon the temperature, and tabulated values are found in reference tables.

Introduction

The molar mass of a gaseous compound can be determined by experiment even though the formula or composition of the compound is not known. In other words, it is possible to find the molar mass of a gas even if we do not know its identity. The molar mass is determined by using the fact that the number of moles of a gas sample is related to the pressure, volume, and temperature by the gas laws. The mass, pressure, volume and temperature of a gas sample are measured experimentally, and the data is used to calculate the molar mass.

To accomplish this, the number of moles in the sample is first calculated using the ideal gas law, $PV = nRT$. Recall that n represents the number of moles. The number of moles of a gas is found algebraically as:
$n = PV/RT$ where R is the universal gas constant.

As an example, consider a 0.508 g sample of a gas that occupies 522 mL at 100 °C and 0.960-atmosphere pressure. First summarize the data as:

Mass	P	V	T
0.508 g	0.960 atm	0.522 L	(100.0°C + 273.2) = 373.2 K

The number of moles of gas is found using the ideal gas law. (Fill in the following blanks and calculate the number of moles.)

$$n = \frac{\underline{\quad}\ \text{atm} \times \underline{\quad}\ \text{L}}{0.0821\ \frac{\text{L atm}}{\text{mole K}} \times \underline{\quad}\ \text{K}} = \underline{\quad}\ \text{mole}$$

The molar mass of the gas is found from the ratio of the mass of the sample to the moles in the sample. Divide the mass of the sample by the number of moles using the value calculated above.

$$\frac{0.508\ \text{g}}{\underline{\quad}\ \text{mole}} = \frac{\underline{\quad}\ \text{g}}{1\ \text{mole}}$$

In this experiment, a sample of gas is collected by water displacement; i.e. the gas bubbles into a container of water and as the gas accumulates, it displaces the water. The volume of the sample is the volume the gas occupies in the container. The temperature is found by measuring the temperature of the water in contact with the gas. The mass of the gas sample is measured by weighing a small tank or cylinder of gas, delivering the sample, and reweighing the cylinder.

The pressure of the sample is found by making sure that the pressure of the gas sample is the same as the atmospheric pressure. The atmospheric pressure is measured with a barometer. However, when a gas is collected by water displacement, it becomes saturated with water vapor. This means that once the gas sample is collected, it will be a mixture of the gas and water vapor. This does not affect the volume or temperature of the gas, but the measured pressure is the total pressure of the gas and the water vapor. To determine the pressure of the gas sample, the pressure of the water vapor must be subtracted from the total pressure. Appendix 2 lists the vapor pressure of water at various temperatures. To find the pressure of the gas sample in the experiment, look up the vapor pressure of water at the measured temperature and subtract this pressure from the measured atmospheric pressure:

$$P_{GAS} = P_{ATM} - P_{WATER}$$

Experimental Procedure

For this experiment you will need a 1-L beaker, a 250-mL Erlenmeyer flask, and a thermometer.
Place a piece of magic tape or a length of a gummed label along the neck of the flask near the top. Fill the beaker to the 750mL mark with tap water and transfer 250mL of this volume into the conical flask, filling it to the brim. Allow the water to stand for several minutes to be sure that it is at room temperature. Set up the apparatus as shown in Figure 1 below.

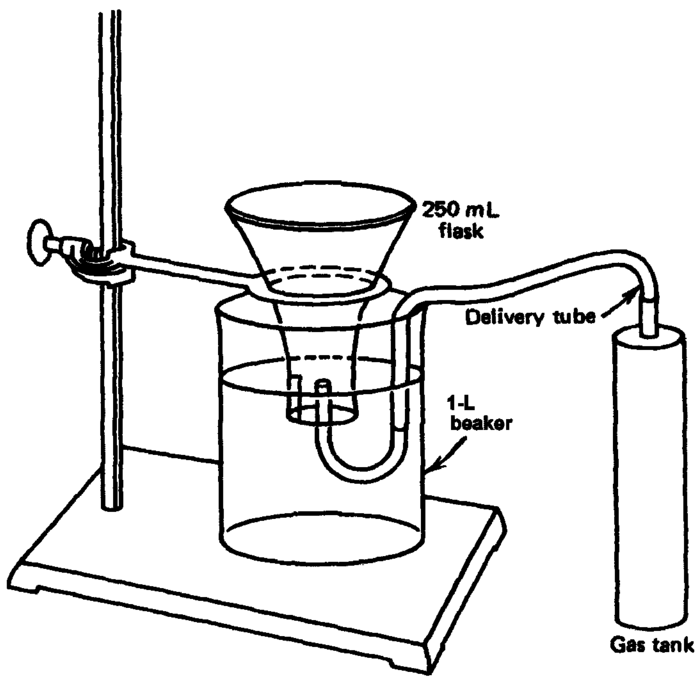

Figure 1. Apparatus for collecting gas samples

Gently stopper the flask with a rubber stopper and do not allow any air to become trapped in the flask under the stopper. Invert the flask through the iron ring that will hold it in position. Lower the ring with the flask so that the mouth of the flask is below the water level in the beaker. Secure the ring to support the flask. Use a spatula to remove the rubber stopper from the mouth of the flask. The stopper can remain in the beaker since it will not interfere with the experiment.

Obtain a small tank of gas. Use a paper towel to gently wipe the tank to make sure that it is clean. Weigh the tank to the nearest 0.01 g.

Attach a plastic delivery tube to the valve on the tank and insert the glass tube on the other end into the mouth of the flask in the beaker of water. Push the valve on the tank straight down to deliver gas to the flask. Continue to deliver gas until the flask is nearly full of gas, i.e. deliver gas until the gas level is near the neck of the flask. Do not allow any gas to escape from the flask. If any gas escapes, you will have to start the experiment over.

Carefully remove the delivery tube from the gas tank, gently clean and dry the tank with a paper towel and reweigh the tank to the nearest 0.01 g. Place a thermometer in the beaker.

Move the flask of gas up or down so that the level of water in the flask is the same as the water level in the beaker. This will adjust the pressure in the flask to atmospheric pressure. Use a pen or pencil to place a mark on the tape to indicate the water level.

Remove the flask from the apparatus and fill it with water to the level corresponding to the mark on the tape. This volume of water will correspond to the volume of the gas sample. Carefully pour the water from the flask into a graduated cylinder and measure the volume to the nearest 1 mL. Record this as the volume of the gas sample.

Record the temperature of the water in the beaker. Assume that the temperature of the gas sample is the same as the water. Record the atmospheric pressure and look up the vapor pressure of water at the temperature of the gas in the table at the end of this experiment. Repeat the experiment to obtain a second set of data. Do all your calculations before returning your equipment.

Calculate the molar mass of the gas using each set of data. First obtain the mass of the gas by subtracting the mass of the can after the sample was removed from the mass of the can before removing the sample. Calculate the pressure of the gas by subtracting the vapor pressure of water from the total atmospheric pressure. The vapor pressure of the water vapor in your gas sample is dependent on temperature and is found in the following table. If your temperatures are not exactly the same as those in the table, calculate the vapor pressure for your temperatures as measured to the tenths place on your lab thermometers. For example, at a temperature of 15.2 ^{o}C, the vapor pressure of water is between 12.8 and 13.6 torr. Since one degree Celsius (16 ^{o}C -15 ^{o}C) in this range equals 0.8 torr (13.6 torr-12.8 torr), 0.2 degree Celsius corresponds to an additional 0.16 torr, so that the vapor pressure of water at 15.2 ^{o}C is 13.0 torr (12.8 torr + 0.16 torr = 12.96 torr, or 13.0 torr).

Table 1. Vapor pressure of water as a function of temperature

Temperature (°C)	**Vapor Pressure (torr)**	**Temperature (°C)**	**Vapor Pressure (torr)**
0	4.6	25	23.8
5	6.5	26	25.2
10	9.2	27	26.7
15	12.8	28	28.3
16	13.6	29	30.0
17	14.5	30	31.8
18	15.5	40	55.3
19	16.5	50	92.5
20	17.5	60	149.4
21	18.6	70	233.7
22	19.8	80	355.1
23	21.2	90	525.8
24	22.4	100	760.0

Again, make sure your calculated answers (A-C) have significant figures that are consistent with the original data. Then convert your gas sample quantities into units that are consistent with the gas constant, R. Use algebra and the ideal gas law to calculate the molar mass. You can solve for the number of moles, n, and then divide the gram quantities by this number to get the molar mass. Another option is to derive the molar mass in units of g/mol directly from ideal gas law. If the results of the two trials are within 10 percent of one another, express your answer as the average of the two results.

If your two results differ by more than ±10%, run a third trial. Consult your instructor to see if you should include all three trials in your average.

Activity 15 – Molar Mass of a Gas

Name____________________________

Section____________ **Date**_______

Experimental Data: Record all your observations.

	Trial 1	Trial 2	Trial 3
Raw Data			
Mass of tank (g)			
Mass of tank – sample (g)			
Mass of sample (g)			
Volume of sample (mL)			
Temperature of sample (°C)			
Atmospheric pressure (torr)			
Vapor pressure of water (torr)			
Pressure of sample (torr)			
Conversion of Data to Appropriate Units			
Mass of sample (g)			
Volume of sample (L)			
Temperature of sample (K)			
Pressure of sample (atm)			
Calculate Values			
Moles of sample (mol)			
Molar Mass of sample (g/mol)			

Show a sample calculation for one trial and calculate the average molar mass:

Average Molar Mass (g/mol) ____________

Questions and Problems

1. Referring to the experimental determination of the molar mass, explain how and why each of the following factors would affect your calculated molar mass. Would the calculated value be greater than it should be, less than it should be, or not changed? Note that the relationships between the various factors involved in the calculation are:

$$\text{Molar mass} = \frac{g}{mol} = \frac{g}{n} = \frac{g}{\left(\frac{PV}{RT}\right)} = \frac{gRT}{PV}$$

   a. The measured temperature is a lower value than the actual temperature.

   b. The measured volume is a higher value than the actual volume.

   c. Some of the gas sample escapes from the tank before it reaches the flask.

2. The molar mass of a gas is determined by collecting a gas sample by water displacement.

   a. Using the following data, calculate the molar mass of the gas: Sample volume, 163 mL; temperature, 21.0 ^{o}C; mass, 0.281 g; total (i.e., barometric) pressure, 752 torr.

   b. The gas in part (a) contains 85.5% C and 14.5% H. First, determine the empirical formula, and then use the result from part (a) to determine the actual formula. First step: Assume 100. gram sample.