**Experiment 14**

# Molar Mass of a Gas

## Purpose

In this experiment you will use the ideal gas law to calculate the molar mass of a volatile liquid compound by measuring the mass, volume, temperature, and pressure of the compound in its gaseous state.

## Background

There are several methods by which the molecular weight of a compound can be determined in the laboratory. In this experiment you will use a method that can be applied to volatile liquids. Volatile liquids are those liquids that easily turn into a gas when heated. Most of these types of compounds will behave like an ideal gas when converted to the vapor state. This means that the **ideal gas law** will apply:

$$\mathbf{PV = nRT}$$

In this equation, **P** is the **pressure** of the gas, **V** is the **volume** of the gas, **n** is the **amount** of the gas in **moles**, and **T** is the **temperature** of the gas in Kelvin. **R** is called the **ideal gas constant**. The value of R will differ depending on the units used for pressure and volume. When P is in atmospheres and V is in liters, the value of R is **0.0821 (L atm)/(mol K)**.

This equation is useful because it allows one to calculate the pressure, volume, temperature or number of moles of a gas simply by knowing the other three variables and doing a little algebra.

In the following experiment you will use a setup with which you can easily determine the values for pressure, volume and temperature of a gas. Once these values have been found, you can determine the amount of the gas in moles, and from the mass of the gas in grams, you can calculate the molar mass (molecular weight) of the gas as follows:

$$\textbf{moles of gas} = \mathbf{n} = \frac{\mathbf{PV}}{\mathbf{RT}}$$

Once the number of moles in a sample are calculated, the molar mass of the gas can be calculated by taking the mass of the gas and dividing by the moles of the gas in the sample

$$\textbf{molecular mass} = \frac{\textbf{grams of}}{\textbf{moles of gas}}$$

**EXAMPLE**. To calculate the molar mass of a volatile liquid, the liquid was vaporized in an Erlenmeyer flask which had a total volume of 152 mL. In the procedure, the flask containing an excess amount of the volatile liquid was covered with aluminum foil with a tiny pinhole, and then the flask and the liquid was placed in a boiling water bath at 100 °C. The atmospheric pressure was measured 754 torr with a barometer in the room.

As the liquid in the flask vaporized, the excess vapor escaped through the pinhole until no visible liquid remained in the flask. The flask was then removed from the water bath and allowed to cool. The mass of the flask, foil, and vapor was 94.53 g. The initial mass of the dry, empty flask and foil was 94.12 g.

From this information, calculate the molar mass of the volatile liquid in grams per mole.

**ANSWER**. The calculation is actually quite straightforward, but we must take care to use the correct units on all numbers. Pressure should be in atmospheres, volume in liters, and of course temperature in Kelvin. The calculation consists of two parts:

*General Steps:*

***1)*** *Convert values in the correct units*

$$P = \frac{754 \text{ torr}}{} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.9921 \text{ atm}$$

$$V = \frac{152 \text{ mL}}{} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.152 \text{ L}$$

$$T = 100 \text{ °C} + 273.15 = 373.15 \text{ K}$$

**2)** *Use PV = nRT to calculate the moles of gas, (n)*

$$n = \frac{(0.9921 \text{ atm})(0.152 \text{ L})}{(0.08206 \text{ L atm / mol K})(373.15 \text{ K})} = 0.00492 \text{ moles}$$

***3)*** *Divide the grams of gas by the moles of gas (calculated in Step 1) to obtain the molecular weight of the gas in grams per mole.*

The mass of the gas = 94.53 g – 94.12 g = 0.41 g

$$\text{The molar mass of the gas} = \frac{0.41 \text{ g}}{0.00492 \text{ mol}} = 83 \text{ g/mol}$$

## Procedure:

**Safety:** Be sure to wear your safety goggles and take care with hot glassware.
You unknown may be flammable so avoid any open flames.

**Waste:** This lab will not generate any hazardous waste.

**Materials**

- 600 mL beaker
- 125 mL Erlenmeyer flask
- 10 mL Cylinder
- 50 or 100 mL Grad. Cylinder
- Thermometer
- Ring Stand
- Clamp or Iron Ring
- Aluminum Foil
- Rubber Bands
- Unknown Liquids

**1**) Determine the combined total mass of a dry 125 mL Erlenmeyer flask, a rubber band, and a square of aluminum foil, and record this mass on the report form. NOTE: If you need to exchange your flask for any reason, be sure to re-weigh the new flask, as the weights of even the same sized flasks vary considerably.

**2**) Obtain a sample of an unknown liquid from your instructor. **Record the unknown number** in the Data Section.

**3**) Create a hot water bath by filling a 600 mL beaker half full of water. Heat to boiling using a Bunsen Burner, and while waiting for the water to boil, use a 10 mL graduated cylinder to measure and pour approximately 2 mL of your liquid unknown sample into the flask.
Secure the aluminum foil over the mouth of the flask with the rubber band. Do not use too big a piece of foil since if it gets wet it can affect your measurements. Make one tiny hole in the foil with a pin to let excess vapor escape during heating.

**4**) Clamp the flask assembly into the beaker so that flask is as far down as possible in the beaker. Heat the flask in boiling point water for 10 minutes. Measure the temperature of the gas occupying the flask by recording the temperature of the boiling water surrounding the flask to the nearest ±0.1 $^{o}$C.

**5)** Since the flask is even so slightly open to the atmosphere, the pressure of the gas must be equal to the barometric pressure. Measure the atmospheric pressure with the barometer in the laboratory and record the current barometric pressure on the report form.

**6**) Remove the flask from the water bath by loosening the clamp from the supporting rod and moving it upward. Allow the flask to cool to room temperature, and dry it on the outside gently and thoroughly with a paper towel. Weigh the flask along with its contents, the aluminum foil and rubber band, and record this value on the report form.

**7**) Accurately measure the volume of the flask (it is more than 125 mL!) by filling it with water all the way to the brim and measuring the volume of water with a 100 mL graduated cylinder. Record the flask volume on the report form.

**8**) Calculate the moles of vapor from its pressure, volume, and temperature, and from this value and the mass of the vapor, calculate the molar mass of the vapor.

**Name** ______________________________

## Data Table  Trial 1

Unknown Number _____________

1) mass of flask, aluminum foil and rubber band __________ g

2) Temperature of boiling water ________ ºC __________ K

3) Atmospheric Pressure _________ mm Hg __________atm

***(show conversion calculation below)***

4) Mass of flask, aluminum foil, rubber band, and condensed vapor ___________ g

5) Mass of condensed vapor __________g

6) volume of flask ________ mL __________L

***(show conversion calculation below)***

7) Moles of vapor (from PV = nRT) ________ mol

***(show your calculation below)***

8) Molar Mass of unknown _______ g/mol

***(show your calculation)***

## Data Table Trial 2

Unknown Number _______________

1) mass of flask, aluminum foil and rubber band ___________ g

2) Temperature of boiling water _________ $^{o}C$ ___________ K

3) Atmospheric Pressure __________ mm Hg `___________atm
***(show conversion calculation below)***

4) Mass of flask, aluminum foil, rubber band, and condensed vapor ____________ g

5) Mass of condensed vapor ___________g

6) volume of flask _________ mL ___________L
***(show conversion calculation below)***

7) Moles of vapor (from PV = nRT) _________ mol
***(show your calculation below)***

8) Molar Mass of unknown ________ g/mol
***(show your calculation below)***

## Post-Lab Questions

1. When the water starts boiling, what happens to the volatile liquid in the flask? Circle **all** that apply.

i. The liquid solidifies due to the higher pressure in the flask.

ii. The liquid evaporates.

iii. The liquid is converted to a gas and it pushes air out of the flask.

2. Why is it not necessary to weigh the flask and the liquid unknown before the initial heating?

3. Calculate the molar mass of a gas if 35.44 g of the gas stored in a 7.50 L tank exerts a pressure of 60.0 atm at a temperature of 35. 5 $^{o}$ C

4. What temperature will 0.654 moles of neon gas occupy 12.30 liters at 1.95 atmospheres?

5. For each of the following errors in the experiment, indicate if the experimentally calculated value of the molar mass would be a higher value, lower value or same value if

a. the experimental mass of the gas in the flask recorded was lower than the true value

b. the experimental temperature recorded was lower than the true value

c. the experimental pressure recorded was higher than the true value

**Name ______________________**

**Pre-Laboratory Assignment for Molar Mass of a Gas**

1. Describe the relationship between the pressure and volume of a gas.

2. Describe the relationship between the pressure and the temperature of a gas.

3. A sample of He(g) is kept in a container with a volume of at 23.8 mL and at 0.971 atm. If the pressure of the gas is decreased to 0.828 atm, will its volume increase or decrease? Calculate the new volume of the gas at this pressure.

4. What does the term "volatile" mean? Why is it necessary for your unknown to be volatile for this experiment?

5. Calculate the molar mass of an unknown gas if 975 mL of it at 21 $^{\circ}$ C and 0.962 atm weights 9.23 g?