Molar Mass of a Volatile Liquid using Ideal Gas Law

INTRODUCTION

The most common instrument for the determination of molar masses in modern chemical research is the **mass spectrometer**. Such an instrument permits very precise determination of molar mass and also gives a great deal of structural information about the molecule being analyzed; this is very helpful in the identification of new or unknown compounds. Mass spectrometers, however, are extremely expensive and take a great deal of time and effort to calibrate and maintain. For this reason, many of the classical methods of molar mass determination are still widely used.

In this experiment, a common modification of the ideal gas law will be applied in the determination of the molar mass of a liquid that is easily evaporated. The ideal gas law ($PV = nRT$) indicates that the observed properties of a gas sample [pressure (P), volume (V), and temperature (T)] are directly related to the quantity of gas in the sample (n, moles). For a given container of fixed volume at a particular temperature and pressure, only one possible quantity of gas can be present in the container:

$$n = \frac{PV}{RT}$$

By careful measurement of the mass of the gas sample under study in the container, the molar mass of the gas sample can be calculated, since molar mass (MM) merely represents the number of grams (m) of the volatile substance per mole:

$$n = \frac{m}{MM}$$

In this experiment, a small amount of easily volatilized liquid will be placed in a flask of known volume. The flask will be heated in a boiling water bath and will be equilibrated with atmospheric pressure. From the volume of the flask used (V), the temperature (T) of the boiling water bath, and the atmospheric pressure (P), the number of moles of gas (n) contained in the flask may be calculated. From the mass of liquid required to fill the flask with vapor when it is in the boiling water bath, the molar mass of the liquid may be calculated.

A major assumption is made in this experiment that may affect your results. We assume that the vapor of the liquid behaves as an *ideal* gas. Actually, a vapor behaves *least* like an ideal gas under conditions near which the vapor would liquefy. The unknown liquids provided in this experiment have been chosen, however, so that the vapor will approach ideal gas behavior.

Name _______________________________________ Date _______ Grade ____

PRE-LAB QUESTIONS

MUST be completed before an experiment is started. The COPY pages will be collected as you enter the lab.

Please answer the following questions and show all work and units. Express all answers to the correct number of significant digits.

1. Write the ideal gas equation; define and specify the unit for each variable.

2. The variable n can be written as

$$\frac{\text{mass of ideal gas}}{\text{molar mass of ideal gas}} \text{ or } \frac{g}{MM}.$$

   Substitute this quotient for n in the ideal gas equation and solve for the molar mass.

3. Describe the experimental data which would be needed to determine the molar mass of an unknown gas.

4. Read the introduction to the experiment on the following page. Explain why it is necessary to remove the Erlenmeyer flask from the hot water bath at the moment the volatile liquid disappears from the flask, rather than while there is still liquid in the flask, or minutes after the liquid has completely evaporated.

EXPERIMENTAL PROCEDURE

SAFETY PRECAUTIONS

• Wear safety goggles at all times while in the laboratory.

• Assume that the vapors of your liquid unknown are toxic. Work in an exhaust hood or other well-ventilated area.

• The unknown may also be flammable. All heating is to be performed using a hot plate.

• The liquid unknowns may be harmful to skin. Avoid contact, and wash immediately if the liquid
is spilled.

• A boiling water bath is used to heat the liquid, and there may be a tendency for the water to splash when the flask containing the unknown liquid is inserted. Exercise caution.

• Use tongs or a towel to protect your hands from hot glassware.

1. Record all data and observations directly in your lab report.
2. Rinse a 125-mL Erlenmeyer flask with tap water then with deionized water. Use about 5 mL acetone to completely dry the flask. The flask must by *completely dry*, since any water present will vaporize under the conditions of the experiment and will adversely affect the results.
3. Cut a square of thick (freezer) aluminum foil to serve as a cover for the flask. Trim the edges of the foil so that it neatly covers the mouth of the flask but does not extend far down the neck.
4. Weigh the dry, empty flask with its foil cover.
5. Obtain an unknown liquid and record its identification number.
6. Add 3 to 4 mL of liquid to the dry flask. Cover the flask with the foil cover, making sure that the foil cover is tightly crimped around the rim of the flask. Punch a single small hole in the foil cover with a needle or pin.
7. Prepare a 600-mL beaker for use as a heating bath for the flask. The beaker must be large enough for most of the flask to be *covered* by boiling water when in the beaker. Add the required quantity of water to the beaker and add a few boiling chips. Set up the beaker on a hot plate close to the exhaust.
8. Heat the water in the beaker to boiling. When the water in the beaker begins to boil, adjust the temperature of the hot plate so that the water remains boiling but does not splash from the beaker.
9. Immerse the flask containing the unknown liquid in the boiling water so that most of the flask is covered with the water of the heating bath (see Figure 1). Clamp the neck of the flask to maintain the flask in the boiling water.

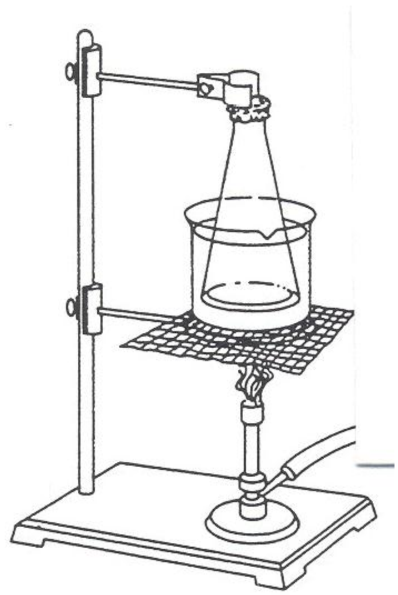

Figure 1. Apparatus for determination of the molar mass of a volatile liquid. Most of the flask containing the unknown liquid must be beneath the surface of the boiling water bath. [**Note: use hot plate instead of bunsen burner.**]

10. Watch the unknown liquid carefully. The liquid will begin to evaporate rapidly, and its volume will decrease. The amount of liquid placed in the flask is much *more* than will be necessary to fill the flask with vapor at the boiling water temperature. Excess vapor will be observed escaping through the pinhole made in the foil cover of the flask.
11. When it appears that all the unknown liquid has vaporized, and the flask is filled with vapor, continue to heat for 1 or 2 more minutes. Then remove the flask from the boiling water bath; use the clamp on the neck of the flask to protect your hands from the heat.
12. Set the flask on the lab bench, remove the clamp, and allow the flask to cool to room temperature. Liquid will *reappear* in the flask as the vapor in the flask cools. While the flask is cooling, measure and record the exact temperature of the boiling water in the beaker, as well as the barometric pressure in the laboratory.
13. When the flask has cooled *completely* to room temperature, carefully dry the outside of the flask to remove any droplets of water. Then weigh the flask, foil cover, and condensed vapor.
14. **Do not wash the flask for the second trial.**
15. Repeat the determination by adding another 3 to 4 mL sample of unknown liquid. Reheat the flask until it is filled with vapor, cool, and reweigh the flask. The weight of the flask after the second sample of unknown liquid is vaporized should agree with the first determination within 0.05 g. If it does not, do a third determination.
16. When two acceptable determinations of the weight of vapor needed to fill the flask have been obtained, remove the foil cover from the flask and clean it out.
17. Fill the flask to the very rim with tap water, cover with the foil cover, and weigh the flask, cover, and water. Determine the temperature of the tap water in the flask. Using the density of water at the temperature of the water in the flask and the weight of water the flask contains, calculate the exact volume of the flask.
18. If no balance with the capacity to weigh the flask when filled with water is available, the volume of the flask may be approximated by pouring the water in the flask into a 1-L graduated cylinder and reading the water level in the cylinder.
19. Using the volume of the flask (in liters), the temperature of the boiling water bath (in

kelvins), and the barometric pressure (in atmospheres), calculate the number of moles of vapor the flask is capable of containing. [R = 0.08206 L atm mol^{-1} K^{-1}].

20. Using the weight of unknown vapor contained in the flask, and the number of moles of vapor present, calculate the molar mass of the unknown liquid.

Clean-Up. Clean all glassware that was used before leaving the laboratory. Place all waste solutions into proper waste containers.

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Name ______________________________ Date ________

Partner's Name __________________________

DATA AND CALCULATIONS

Identification number/letter of unknown liquid _____________

Mass of empty flask and cover _______________ g

	Sample 1	**Sample 2**
Mass of flask+cover+vapor		
Temperature of vapor (°C)		
Temperature of vapor (K)		
Pressure of vapor (mm Hg)		
Pressure of vapor (atm)		
Mass of flask+cover+water		
Mass of water in flask		
Temperature of water in flask		
Density of water		
Volume of flask		
Moles of vapor in flask		
Molar mass of vapor		
Mean value of molar mass of vapor		
Density of vapor in flask		
Mean value of density of vapor		

Show sample calculations:

Post-Lab Questions and Exercises

(All questions must be answered during the lab and submitted with your lab report at the end of the lab period).

Please answer the following questions and show all work and units. Express all answers to the correct number of significant digits.

Q1. If 2.31 g of the vapor of a volatile

Q2. liquid is able to fill a 498-mL flask at 100°C and 775 mm Hg, what is the molecular weight of the liquid? What is the density of the vapor under these conditions?

Q3. Why is a vapor unlikely to behave as an ideal gas near the temperature at which the vapor would liquefy?

Q4. Two methods were described for determining the volume of the flask used in the molar mass determination. Which method would give a more precise determination of the volume? Why?

Q5. Use the ideal gas law to determine the molecular weight of the unknown. If you completed more than one experimental run, calculate the average value of the molecular weight. Show your work!

Q6. Complete the table for each sample run - Show any calculations including unit conversions.

Pressure	Volume	Mass of Vapor	Ideal gas constant	Temperature
atm	L	g	$\frac{L\cdot atm}{mol\cdot K}$	K
atm	L	g	$\frac{L\cdot atm}{mol\cdot K}$	K
atm	L	g	$\frac{L\cdot atm}{mol\cdot K}$	K

Q7. Show your work to your instructor. If the calculation is correct he/she will initial it and give you the name of your unknown. From its formula, calculate the molecular weight of the unknown.

Q8. Calculate the percent error of your measured molecular weight.

$$\%\ \text{error} = \frac{|\text{accepted value - experimental value}|}{\text{accepted value}} \times 100$$

Q9. A student failed to vaporize the entire sample prior to placing the test tube in the ice bath. How did this error affect the calculated molar mass? Justify your answer using calculations.

Q10. A different student failed to dry the outside of the test tube prior to massing it. How did this error affect the calculated molar mass? Justify your answer using calculations.

Q11. How would your calculated molar mass have been affected if you had used twice the initial amount of the unknown compound?