DETERMINING THE MOLAR MASS
OF A VOLATILE LIQUID

OBJECTIVES

To use the physical properties of a gas phase sample and the ideal gas law to determine the molar mass of a volatile liquid

MATERIALS

One 125-mLErlenmeyer flask; one 600-mL beaker; one 25-mL graduated cylinder; one 50- or 100-mL graduated cylinder; aluminum foil; ring stand with two cast-iron support rings, clamp, and wire gauze pad; Bunsen burner or hot plate; analytical balance; thermometer; boiling chips; samples of unknown volatile liquids

SAFETY

Most volatile liquids are also flammable. Take appropriate precautions when heating volatile liquids near open flames. Safety goggles should be worn at all times

WASTE DISPOSAL

All recovered or unused samples of volatile liquids should be collected and disposed in an appropriate Organic Waste container.

INTRODUCTION

The determination of **molar mass** represents an important step in the identification of an unknown substance. There are many methods to obtain this vital information that are related to physical properties (vapor pressure, osmotic pressure), while others rely on chemical behavior, such as reactions of known stoichiometry. In this lab you will determine the molar mass of an unknown volatile liquid by taking advantage of the properties and behavior of gaseous substances, and the ideal gas law.

The physical properties of matter in the solid and liquid states (i.e., melting points, boiling points) are heavily dependent on the intermolecular forces between molecules. Strong attractive forces, such as dipole-dipole and hydrogen bonding), lead to significant increases in melting points and boiling points. Likewise, increases in the size of molecules results in increases in dispersive forces, and a corresponding increase in the observed phase transition temperatures.

By contrast, the behavior of matter in the vapor state is uniformly predictable using the various gas laws. The volume occupied by gases increases with increasing temperature, decreases with increasing pressure, and increases as the number of moles of matter increases. The uniformity of behavior of gases is explained by the postulates of the **Kinetic Molecular Theory of Ideal Gases**, which are summarized here:

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1. Molecules in the vapor state are separated by large distances, so interactions between molecules are negligible.

2. Collisions between molecules in the vapor state are elastic; kinetic energy is conserved.

3. The volume occupied by the molecules themselves is negligible compared to the total volume occupied by the gas.

An important consequence of the kinetic molecular theory is that all gases behave similarly, regardless of the identity of the substance. Overall, the behavior of gases can be summarized by the **Ideal Gas Law**,

$$PV = nRT \qquad (1)$$

where P = pressure (in atm), V = volume (in L), n = moles of gas, T is the temperature (in kelvins), and R is the universal gas constant (0.0821 L·atm/ mol·K). If the volume occupied by a gas sample is known at a given temperature and pressure, then the Idea Gas Law can be rearranged to solve for n:

$$n = \frac{PV}{RT} \qquad (2)$$

The mass of a vapor sample can be measured as the difference between the mass of an empty container and the mass of that same container with the vapor sample. Once the mass is known, and the number of moles of the vapor sample has been calculated using Equation (1), then the molar mass of the substance can be calculated:

$$M_{compound} = \frac{mass_{sample}}{n_{sample}} \qquad (3)$$

These uniformity of these relationships provides a convenient way to determine the molar mass of many substances, especially liquids, as long as they are volatile enough to readily enter the vapor state under relatively moderate conditions of temperature and pressure.

Although the Ideal Gas Law works well for most gases under ambient conditions, the assumptions inherent in the kinetic molecular theory are no longer valid at extreme conditions of high pressure and low temperature. Under these conditions, molecular collisions are less likely to be elastic and more likely to be influenced by attractive intermolecular forces between molecules. The Van der Waals equation, a modification of the Ideal Gas Law, corrects for these deviations from ideality:

$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT \qquad (4)$$

The P, V, n, R, and T terms are the same as defined previously. The a term corrects for attractive forces between molecules while the b term corrects for the actual volume occupied by the gas molecules themselves. Physical data, including Van der Waals constants, for several volatile organic liquids are included in Table 1.

Table 1. Physical Data for Volatile Organic Liquids.

Compound	Formula	Molar Mass	a ($L^2 \cdot atm/mol^2$)	b (L/mol)	b. p. (ºC)
Methanol	CH_3OH	32.04	9.52	0.0670	65.0
Ethanol	CH_3CH_2OH	46.07	12.02	0.0841	78.5
Propanol	$CH_3CH_2CH_2OH$	60.09	14.92	0.1019	82.4
Acetone	CH_3COCH_3	58.02	13.91	0.0994	56.5
Pentane	$CH_3(CH_2)_3CH_3$	72.14	19.01	0.1460	36.0
Hexane	$CH_3(CH_2)_4CH_3$	86.17	24.39	0.1735	69.0
Cyclohexane	$C_{-6}H_{12}$	84.15	22.81	0.1424	80.7
Water	H_2O	18.02	5.46	0.0305	100.0

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PRE – LAB QUESTIONS

1. A student performing this experiment records a volume of 175 mL at a barometric pressure of 755 torr and 93.C. Using the Ideal Gas Law, calculate the number of moles of vapor in the flask under these conditions. Let R = 0.0821 L·atm/mol·K, and perform all necessary conversions before calculating.

2. A student performs this experiment and obtains the data presented in Question 1, but forgets to record the temperature of the water bath and uses a temperature of 298 K when calculating the number of moles. Calculate the value of n he or she would obtain, and the % error that would result.

3. List three postulates of the kinetic molecular theory inherent in the ideal gas law.

PROCEDURE

**Part A**: Determining Volume of Container

1. Obtain a clean, dry 125-mL Erlenmeyer flask, and fill it to the brim with water.

2. Carefully transfer the water from the flask to a 50- or 100-mL graduated cylinder. This will require multiple transfers and measurements; record all individual volumes transferred, and sum them to obtain the total volume of the flask.

3. Record the total volume of the 125-mL flask on your Data Sheet.

4. Repeat steps 1–3, and record this total volume on your Data Sheet.

5. Calculate and record the average total volume.

Part B: Determining the Mass of a Volatile Liquid Sample

6. Obtain a 600-mL beaker and add about 300 mL of DI water. Add a few boiling chips to the beaker.

7. Assemble the ring stand with support rings and clamps, wire gauze pad, 600-mL beaker, thermometer, and Bunsen burner or hot plate, as illustrated in Figure 1. Start heating the water in the beaker while completing steps 8–12. Remove from heat if it starts to boil before you reach step 13.

8. Clean and dry the 125-mL Erlenmeyer flask from Part A. Dry the flask either in an oven or by allowing it to air-dry; do *not* wipe it dry with a paper towel, as this may transfer paper residue to the flask and affect the total mass.

9. Cover the flask with a small piece of aluminum foil (about 3-inch square). Secure the foil with a rubber band.

10. Determine the mass of the dry flask, foil, and rubber band and record the mass on your Data Sheet (to the nearest 0.001 g).

11. Using your 25-mL graduated cylinder, obtain about 20 mL of the unknown volatile liquid as assigned by your instructor. If more than one unknown is being used, record the ID for your unknown on your Data Sheet.

12. Carefully remove the rubber band and the foil, and place about 5 mL of your volatile liquid sample in the flask. Replace the foil over the opening of the flask and secure it with the rubber band. Poke three or four pin holes in the foil cap, taking care not to tear the foil.

13. Carefully lower the flask containing the volatile liquid sample into the bath and secure it with a clamp. Adjust the position of the beaker so that the water level is high on the

neck of the flask but not touching the foil cover, and so that the flask is not in contact with the sides of the beaker.

Figure 1. Apparatus for heating volatile liquids
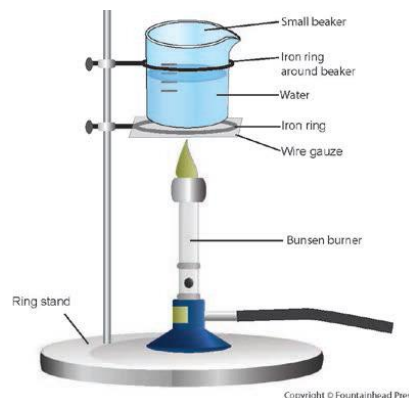
14. Continue heating the water bath while monitoring the state of the volatile liquid sample. You may see vapors escaping though the pin holes while the sample evaporates. When the liquid in the flask is no longer visible, continue heating for another three or four minutes, or until there is no more evidence of vapors escaping from the flask. Record the temperature of the water bath at this point.

15. Remove the flask from the water bath and allow it to cool to room temperature. Some of the vapor in the flask may condense at this point, but this will not affect your results. Dry the outside of the flask and obtain the mass of the flask with the sample, foil cap, and rubber band. Be sure to use the same analytical balance that you used in step 10. Record this mass on your Data Sheet.

16. Perform two additional determinations, repeating steps 12–15 and recording your data for each determination on the Data Sheet.

17. Record the ambient pressure in the lab.

CALCULATIONS

1. Calculate the mass of the vapor in the flask as the difference between the mass of flask-plus sample (recorded in step 15) and the initial mass of the flask with foil cap and rubber band (recorded in step 10). Record this on the Data Sheet.

2. Using the temperature of the water bath (step 14) the ambient pressure (step 17), and the volume of the flask (step 2), calculate the moles of vapor in the flask for each trial using the ideal gas law. Record this on the Data Sheet.

3. From the mass of vapor and the moles of vapor, calculate the molar mass of the vapor and record it on your Data Sheet for each trial.

4. Calculate the average molar mass for your volatile liquid and record it on your Data Sheet.

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DATA SHEET

ID of Unknown Volatile Liquid: _____

Part A. Volume of Flask _____

Trial 1. Volumes: _____

Total volume (1) = _____ (mL)

Trial 2. Volumes: _____

Total volume (2) = _____ (mL)

Average Volume of flask = _____ (mL)

Part B: Determining the Mass of a Volatile Liquid Sample _____

	Trial 1	*Trial 2*	*Trial 3*
Mass of dry flask, foil cap and rubber band:	_____	_____	_____
Mass of dry flask + sample, foil cap and rubber band:	_____	_____	_____
Temperature of water bath:	_____	_____	_____
Barometric pressure:	_____	_____	_____

Calculations:

	Trial 1	*Trial 2*	*Trial 3*
Mass of vapor	_____	_____	_____
Moles of vapor:	_____	_____	_____
Molar Mass:	_____	_____	_____
Average Molar Mass (g/mol):	_____	_____	_____
Actual Molar Mass (g/mol):	_____		
Percent Error	_____		

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POST – LAB QUESTIONS

1. In this lab, the volume of the flask was determined by direct measurement of the volume of water it contained, using a graduated cylinder. Can you propose an alternative—and more precise—method to determine the volume of water contained in the flask?

2. A student performing this experiment accidently tore the foil cap, so that extra vapor escaped from the flask before it was removed from the heat. Describe how this would affect the mass of vapor and the calculated molar mass.

3. A student determines the molar mass of a volatile liquid from the following data using the ideal gas law: V= 138.2 mL, P = 754 torr, T = 371 K, m = 0.330 g. Calculate the molar mass and identify the volatile liquid from the list in Table 1.

4. The Van der Waals equation for a nonideal gas corrects for behavior of gases that do not conform to the postulates of the kinetic molecular theory. Explain the role of the constant a in the Van der Waals equation, and how this also explains the observed trend in a values in Table 1.