Lab #7: Vapour Pressure and Enthalpy of Vaporization of Water

Objectives:

1. Determine the Molar Enthalpy of Vaporization of Water using the Clausius-Clapeyron equation.

Materials:

Thermometer	Barometer
Graduated Cylinder (10 mL)	Ruler
Beaker (1000 mL)	Stirrer
Ice	Ring Stand, Ring, Wire Gauze, Bunsen Burner or Hot Plate

Background Information:

This experiment is designed to find the vapour pressure of water at temperatures between 50°C and 80°C. A graph of the logarithm of vapour pressure versus the reciprocal of absolute temperature allows the calculation of the enthalpy of vaporization.

A sample of air is trapped in an inverted 10 mL graduated cylinder, which is immersed in a tall beaker of water (see Figure 1). As the water in the beaker is heated to about 80°C, the air in the graduated cylinder expands and becomes saturated with water vapour. The temperature and volume are recorded. The total air and water vapour pressure inside the cylinder is equal to the barometric pressure plus a small correction for the pressure exerted by the depth of the water above the trapped air. The water in the beaker is allowed to cool. The volume of air contracts, and less water vapour is present at the lower temperature. The temperature and volume are recorded every 5°C between 80°C and 50°C. Next, the beaker is cooled with ice to a temperature close to 0°C. At this temperature, the vapour pressure if water is so low that it can be assumed that all of the gas in the graduated cylinder is air.

The moles of air molecules in the cylinder can be found by using the volume of dry air present at the temperature near 0°C and the ideal gas equation. Knowing the moles of air in the container, the partial pressure of air can be calculated at each temperature, and the vapour pressure of the water can be obtained by subtracting the pressure of air from the total pressure inside the cylinder.

The Clausius-Clapeyron equation is a mathematical expression relating the variation of vapour pressure to the temperature of a liquid. It can be written:

$$\boldsymbol{\ln P = -\frac{\Delta H_{vap}}{RT} + C}$$

where $\ln P$ is the natural logarithm of the water vapour pressure, ΔH_{vap} is the molar enthalpy of vaporization of water, R is the gas constant (8.314 J/(mol • K)), T is the temperature in Kelvin, and C is a constant which does not need to be evaluated. It can be seen that this equation fits the straight line equation $y = mx + b$ where $y = \ln P$, $x = \frac{1}{T}$, and the slope, $m = -\frac{\Delta H_{vap}}{R}$

If a graph is made of $\ln P$ versus $\frac{1}{T}$, the heat of vaporization can be calculated from the slop of the line.

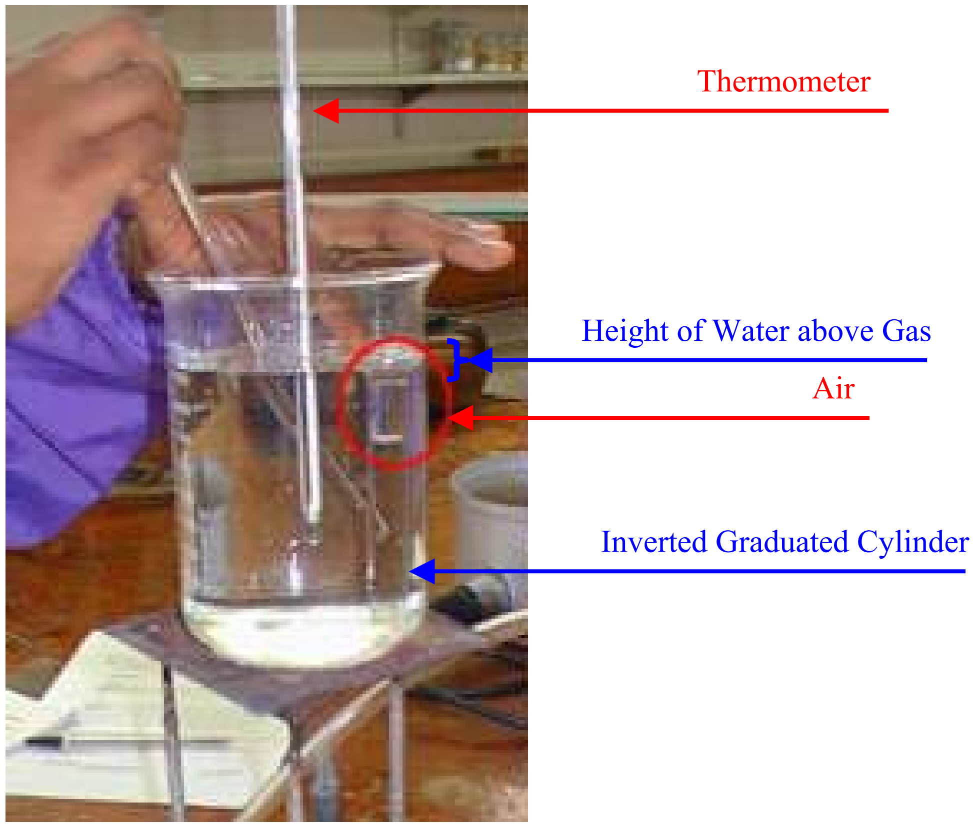

Figure 1: Apparatus Setup for Experiment to determine Molar Enthalpy of Water

Procedure:

1. Put on your lab apron and safety goggles.
2. Obtain a 10 mL graduated cylinder and fill it with approximately 7 mL of water.
3. Fill a 1000 mL beaker approximately three-fourths full with water.
4. Cover the top of the graduated cylinder with your finger and invert it into the 1000 mL beaker. Don't release your finger until the mouth of the graduated cylinder is under the surface of the water in the beaker.
5. If the graduated cylinder is not yet covered with water, add water to the 1000 mL beaker to cover it.
6. Use a ruler and measure the difference between the height of the water in the graduated cylinder and the height of the water in the beaker. This will provide a slight adjustment for the pressure that the water exerts on the air in the graduated cylinder.
7. Record the barometric pressure in mmHg.
8. Heat the 1000 mL beaker on the Bunsen burner until the water is about 80°C. If the air in the cylinder expands past the scale on the graduated cylinder, the experiment will have to be restarted using a smaller amount of air.
9. Record the temperature (±0.1°C) and the volume (±0.01 mL).
10. Cool the beaker until the temperature reaches 50°C recording the temperature and volume every 5°C. Use the same precision as in step 9.
11. Once the beaker reaches 50°C, cool the beaker rapidly to about 0°C by adding ice. Record the gas volume and temperature at this low temperature.

Observations:

Atmospheric Pressure	*mmHg*
Height of Water above Gas	*mmH_2O*

Temperature	Volume

Temperature	Volume

(near 0°C)

Analysis:

1. Graduated cylinders are designed to measure volume standing in the upright position. Because volume is being measured with the graduated cylinder in an inverted position, there will be a slight error, since the meniscus of the water is inverted relative to the graduated cylinder. To account for this error, subtract 0.20 mL from each volume measurement.

Temperature	Volume	Adjusted Volume

Temperature	Volume	Adjusted Volume

(near 0°C)

2. Determine the total pressure in the cylinder by accounting for the pressure exerted by the water that lies above the level of the water in the cylinder.

$$P_{\text{cylinder}} = P_{\text{atmosphere}} + P_{\text{exerted by the water}}$$

The atmospheric pressure will be measured in mmHg, so it will be helpful to convert the amount of water exerting pressure on the gas into mmHg. Water has a density of 1.0 g/mL while mercury has a density of 13.6 g/mL. The pressure exerted by the water (in mmHg) can be determined by simply converting:

$$P_{\text{exerted by water}} = \text{mm } H_2O \times \frac{1.00 \text{ mmHg}}{13.6 \text{ mm } H_2O}$$

Use the above equations, determine the total pressure of the gas in the cylinder.

3. Using the ideal gas law, $PV = nRT$, (R = 62.4 (mmHg • L)/(mol • K)) calculate the number of moles of air in the graduated cylinder at 0°C. The vapor pressure of water at this temperature is so low it is considered negligible, so the pressure of the gas in the cylinder can be considered to be the pressure of the air.

4. With the number of moles calculated in #3, use the ideal gas law to calculate the partial pressure of

air (mmHg) in the cylinder at each of the readings between 50°C and 80°C. Show all calculations for each temperature and adjusted volume used.

Temperature	Partial Pressure of Air (mmHg)

Temperature	Partial Pressure of Air (mmHg)

5. Dalton's Law of Partial Pressures states that the total pressure of a mixture of gases will be equal to the sum of the partial pressures of the component gases. The components of gas in the cylinder that being considered are air (a mixture by itself) and water vapor. The partial pressure of the air is known from #4. Use Dalton's Law of Partial Pressures to calculate the vapor pressure of water at each temperature.

$$\boldsymbol{P_{\text{total}} = P_{\text{air}} + P_{\text{water vapour}}}$$

Temperature	Partial Pressure of H_2O (mmHg)

Temperature	Partial Pressure of H_2O (mmHg)

6. Calculate and enter the values for $\ln P_{\text{water}}$ and $\frac{1}{T}$ for each of temperature measured in the data table below. Next, plot these values on a graph paper with $\ln P_{\text{water}}$ versus $\frac{1}{T}$. Draw the line-of-best-fit through the data points and state its equation along with the correlation coefficient, r (use the Linear Regression and Diagnostic On functions of the TI-83 Plus calculator). Calculate the value of ΔH_{vap} of water using the slope of the line.

Evaluation:

1. The theoretical ΔH_{vap} of water is 40.65 kJ/mol. What is the % error for this experiment?

2. The theoretical ΔH_{vap} for hexane is 28.88 kJ/mol. Explain the difference between it and the ΔH_{vap} of water that was determined in this experiment.

Conclusion:

1. Summarize what you have learned from this lab.