Heat of Vaporization of Ethanol

Equipment

You will need a LabQuest2 with a gas pressure probe (including syringe, rubber stopper assembly, and plastic tubing with two connectors), and a temperature probe. You will also need one 125 mL Erlenmeyer flasks, a 400 mL beaker, a 600 mL beaker, a spoon, and a hot plate.

Chemicals

You will need about 10 mL of 100% pure ethanol.

Introduction

When a liquid is placed in a container, and the container is sealed tightly, a portion of the liquid will evaporate. The newly formed gas molecules exert pressure in the container, while some of the gas condenses back into the liquid state. If the temperature inside the container is held constant, then at some point equilibrium will be reached. At equilibrium, the rate of condensation is equal to the rate of evaporation. The pressure at equilibrium is called *vapor pressure*, and will remain constant as long as the temperature in the container does not change.

In mathematical terms, the relationship between the vapor pressure of a liquid and temperature is described in the Clausius-Clayperon equation,

$$\ln(\mathrm{P}) = -\frac{\Delta \mathrm{H}_{\mathrm{vap}}}{\mathrm{R}}\left(\frac{1}{\mathrm{T}}\right) + \mathrm{C}$$

where ln(P) is the natural logarithm of the vapor pressure, ΔH_{vap} is the heat of vaporization, R is the universal gas constant (8.3145 J/mol·K), T is the absolute, or Kelvin, temperature, and C is a constant. Thus, the Clausius-Clayperon equation not only describes how vapor pressure is affected by temperature, but it relates these factors to the heat of vaporization of a liquid. ΔH_{vap} is the amount of energy required to cause the evaporation of one mole of liquid at constant pressure.

In this experiment, you will introduce a specific volume of a volatile liquid into a closed vessel, and measure the pressure in the vessel at several different temperatures. By analyzing your measurements, you will be able to calculate the ΔH_{vap} of the liquid.

Safety and Waste Disposal

Wear goggles for the entire experiment. **CAUTION:** The alcohol used in this experiment is flammable and poisonous. Avoid inhaling the vapors. Avoid contact with your skin or clothing. Be sure that there are no open flames in the room during the experiment. Notify either your instructor or Matt immediately if an accident occurs.

Procedure

1.) Use a hot plate to heat (boil) about 300 mL of water in a 400 mL beaker.

2.) Prepare a room temperature water bath in a 600 mL beaker. The bath should be deep enough to completely cover the gas level in the 125 mL Erlenmeyer flask.

3.) Use the clear tubing to connect the white rubber stopper to the Gas Pressure Sensor. (About one-half turn of the fittings will secure the tubing tightly.) Twist the white stopper snugly into the neck of the Erlenmeyer flask to avoid losing any of the gas that will be produced as the liquid evaporates (see Figure 1).

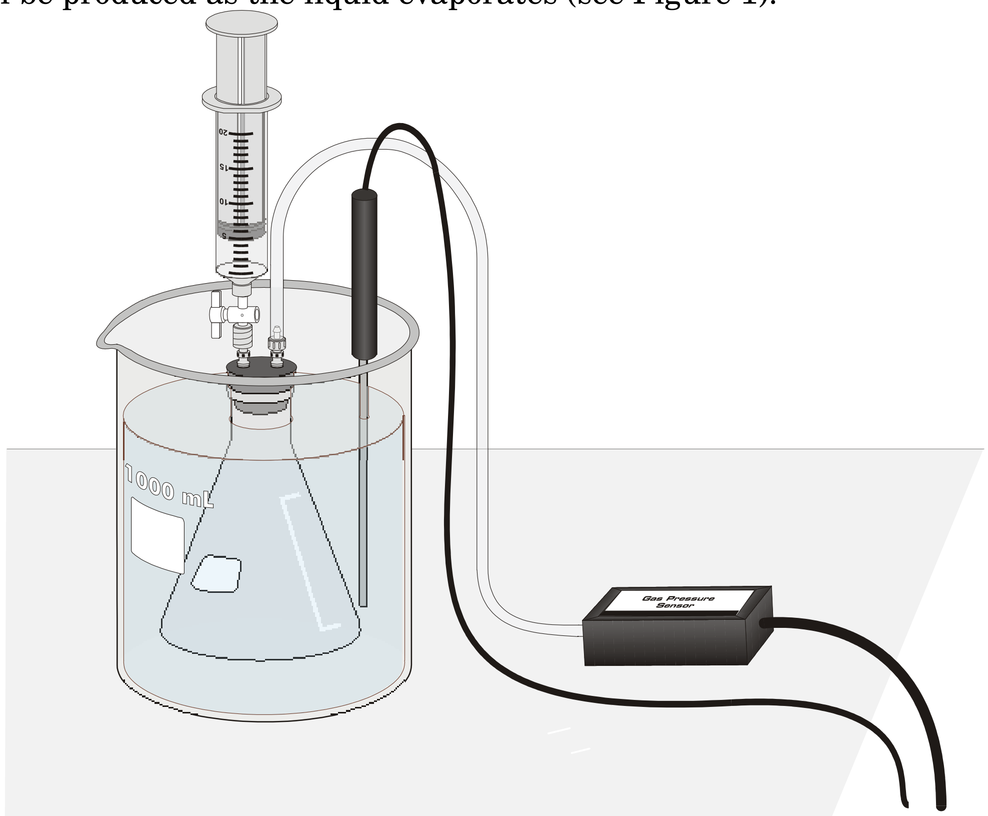

Figure 1

Important: **Open the valve on the white stopper**.

5.) Connect a gas pressure sensor to Channel 1 and a temperature probe to Channel 2 of the LabQuest. Start LoggerPro.

6.) Condition the Erlenmeyer flask and the sensors to the water bath.

a.) Place the Temperature Probe in the room temperature water bath.

b.) Place the Erlenmeyer flask in the water bath. Hold the flask down into the water bath to the bottom of the white stopper.

c.) After 30 seconds, close the valve on the white stopper.

d.) Record the pressure and temperature readings once they stabilize. This is the room pressure and temperature.

d.) Remove the Erlenmeyer flask from the water bath.

7.) Obtain a small amount of ethanol. Draw 3 mL of ethanol into the 20 mL syringe. Thread the syringe onto the two-way valve on the white stopper (see Figure 1).

8.) Add ethanol to the flask.

a.) Open the valve below the syringe containing the 3 mL of ethanol.

b.) Push down on the plunger of the syringe to inject the ethanol.

c.) Quickly pull the plunger back to the 3-mL mark. Close the valve below the syringe.

NOTE: Once you close the valve on the white stopper, do not open it again until step 14!

d.) Carefully remove the syringe from the stopper so that the stopper is not moved.

e.) Place the Erlenmeyer flask back in the water bath.

9.) Gently rotate the flask in the water bath for a few seconds, using a motion similar to slowly stirring a cup of coffee or tea, to accelerate the evaporation of the ethanol.

10.) Monitor and collect the pressure and temperature readings.

a.) Hold the flask steady once again.

b.) Monitor the pressure and temperature readings.

c.) When the readings stabilize, record them.

11.) Remove the Erlenmeyer flask from the 600 mL beaker and add a small amount of hot water, from the beaker on the hot plate, to warm the water bath by about 5°C. Use a spoon to transfer the hot water. Stir the water bath slowly with the Temperature Probe. Place the Erlenmeyer flask back into the 600 mL beaker and monitor the pressure and temperature readings. When the readings stabilize, record them.

12.) Repeat step 11.) until the temperature of the water in the 600 mL beaker is as close to 40 ^{o}C as you can get it without going over 40 ^{o}C. Add enough hot water for each trial so that the temperature of the water bath increases by about 5°C, but **do not warm the water bath beyond 40°C** because the pressure increase may pop the stopper out of the flask. You may have to remove some of the water in the 600 mL beaker so that it does not overflow.

13.) Remove the Erlenmeyer flask and temperature probe from the 600 mL beaker. Pour the water in the 600 mL beaker down the drain and refill it with tap water. Remove the white stopper from the Erlenmeyer flask.

14.) Open the valve on the white stopper and empty the ethanol from the flask into a waste beaker.

15.) Repeat steps 6 through 14 two more times for a total of three run.

Calculations

The P_{air} must be calculated for each different temperature. As you warmed the flask, the air in the flask exerted pressure that you must calculate. Use the ideal gas law (the combined gas law form). Remember that all gas law calculations require Kelvin temperature.

Use the P_{air} from your very first reading (the one with no ethanol in the flask, step 6) as P_1 and the Kelvin temperature for that measurement as T_1. Calculate P_2 at each temperature of that trial. Here P_2 is P_{air} at T_2, which is the temperature in Kelvin.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Calculate the vapor pressure of the ethanol for each trial by subtracting P_{air} from P_{total}. Note each time you recorded a pressure with ethanol in the flask it was P_{total}.

Graph $\ln(P_{vap})$ vs. $1/T(K)$ for each run using a spreadsheet. Have your spreadsheet plot the best-fit straight line, along with the equation. Calculate ΔH_{vap} from the slope of the line.

Remember: $\text{Slope} = \frac{-\Delta H_{vap}}{R}$ so multiply your slope by –R to get the heat of vaporization for that run.

There will be 3 graphs with equations. Include all 3 in your lab report, in the calculations section.

Calculate your average heat of vaporization for ethanol.

The accepted value for the heat of vaporization of ethanol is 42.32 kJ/mol. Calculate your percent error using your average heat of vaporization as the experimental value.

Conclusion

Report your average experimental heat of vaporization for ethanol and your percent error.

Determine and analyze one source of potential experimental error. Please read "How to Determine and Analyze a Source of Experimental Error", the first link in the Laboratory Experiments module on Canvas.