

## EXPERIMENT B2: ENTHALPY OF VAPORIZATION

### Learning Outcomes

Upon completion of this lab, the student will be able to:

- 1) Measure the pressure inside a sealed vessel containing a volatile liquid over a range of temperatures.
- 2) Determine the relationship between pressure and temperature of the volatile liquid.
- 3) Calculate the heat of vaporization of the liquid.

### 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 of the liquid at equilibrium is called the 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:

$$\ln P = -\frac{\Delta H_{vap}}{R} \left( \frac{1}{T} \right) + \text{constant}$$

In the above equation,  $\ln P$  is the natural logarithm of the vapor pressure (in kPa),  $\Delta H_{vap}$  is the heat of vaporization (in kJ/mol),  $R$  is the universal gas constant (8.31 J/mol•K),  $T$  is the absolute, or Kelvin, temperature.

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.

$\Delta H_{vap}$  is the amount of energy required to evaporate one mole of liquid at constant pressure.

## Experimental Design

In this experiment, the enthalpy of vaporization,  $\Delta H_{\text{vap}}$  of ethanol will be determined. Ethanol will be placed inside a tightly sealed container and the vapor pressure of ethanol will be measured as a function of temperature. A plot of  $\ln P$  vs.  $1/T$  will be analyzed to obtain the enthalpy of vaporization.

The sealed container not only contains vapors of ethanol but also the air present inside it. Therefore the measured value of the pressure, which would be the total pressure ( $P_{\text{Total}}$ ), is the sum of the vapor pressure of the ethanol ( $P_{\text{vap}}$ ) and the partial pressure of air ( $P_{\text{air}}$ ) [NOTE: Dalton's law of partial pressures was previously discussed in Experiment B1 (Molar volume of a gas)].

$$P_{\text{Total}} = P_{\text{air}} + P_{\text{vap}}$$

## Reagents and Supplies

Ethanol

(See posted Material Safety Data Sheets)

Vernier LabPro kit with a gas pressure sensor and temperature probe, hot plate (stockroom) laptop computer (lab)

## Procedure

1. Obtain a hot plate, a 600-mL beaker, and Vernier kit with a gas pressure sensor from the stockroom.
2. Obtain a laptop computer from the instructor.
3. Use the hot plate to heat about 200 mL of water in a 600 mL beaker or obtain about 200 mL of hot water from the faucet.
4. Prepare a room temperature water bath in a 600-mL beaker. The bath should be deep enough to reach the bottom of the white stopper, once the apparatus is inserted into the water bath.
5. Setting up the equipment:
  - a. Connect the gas pressure sensor to channel 1 of the LabPro interface box. Connect the temperature probe to channel 2 of the LabPro interface box.
  - b. Connect the LabPro interface box to the power outlet with the power cord. The LabPro will beep to indicate it is ready.
  - c. Connect the LabPro interface box to the laptop computer using the USB cable.
  - d. Connect the computer to the Ethernet using the Ethernet cable and the port located on the lab bench.
6. Prepare the computer to monitor temperature and pressure by opening the file "Vapor Pressure " from the Chemistry 1A desktop folder.
7. 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).  
**Important:** Open the valve on the white stopper.

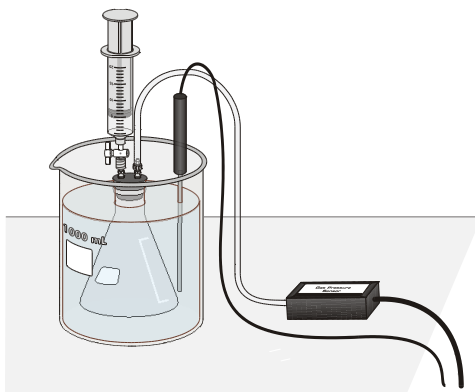


FIGURE 1

8. 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 so that the water level in the beaker is just below the bottom of the white stopper.
  - c. After 30 seconds, close the valve on the white stopper.
  - d. When the pressure and temperature readings stabilize, record these values in the data table as Initial air pressure ( $IP_{\text{air}}$ ) and Initial temperature ( $IT_{\text{air}}$ ).
9. Draw 3 mL of ethanol into the 20 mL syringe that is part of the Gas Pressure Sensor accessories directly from the stock bottle provided. Thread the syringe onto the valve on the white stopper (see Figure 1).
10. 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.
11. 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.
12. Monitor and collect temperature and pressure data.

- a. Click  to begin data collection.
  - b. Hold the flask steady once again.
  - c. Monitor the pressure and temperature readings.
  - d. When the readings stabilize, click .
13. Using beaker tongs add a small amount of hot water, from the beaker on the hot plate, to warm the water bath by 3-5°C. Stir the water bath slowly with the Temperature Probe. Monitor the pressure and temperature readings. When the readings stabilize, click .
14. Repeat Step 13 until a total of five trials have been completed. Add enough hot water for each trial so that the temperature of the water bath increases by 3-5°C, but **do not warm the water bath beyond 40°C** because the pressure increase may pop the stopper out of the flask. If water in the bath must be removed, then do it carefully so as not to disturb the flask.
15. After the fifth set of readings has been recorded, open the valve to release the pressure in the flask. Remove the flask from the water bath and take the stopper off the flask. Dispose the ethanol as directed by the instructor.
16. Click  to end the data collection. Record the first pressure reading,  $P_{\text{total}}$ , and the first temperature reading as Trial 1 in the data table. Record the additional pressure and temperature readings as Trials 2-5.
17. Do not exit the Logger Pro program until steps 1-4 of the data analysis section have been completed.
18. On the second lab period repeat the experiment to collect a second set of data.

## Data Table

### DAY 1

Initial air pressure  $IP_{\text{air}}$  (from Step 8d) = \_\_\_\_\_

Initial temperature  $IT_{\text{air}}$  (from Step 8d) = \_\_\_\_\_

	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
$P_{\text{total}}$ (kPa)					
Temperature (°C)					

### DAY 2

Initial air pressure  $IP_{\text{air}}$  (from Step 8d) = \_\_\_\_\_

Initial temperature  $IT_{\text{air}}$  (from Step 8d) = \_\_\_\_\_

	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
$P_{\text{total}}$ (kPa)					
Temperature (°C)					

## Data Analysis

1. The total pressure inside the flask, which is the measured pressure  $P_{\text{Total}}$ .

$$P_{\text{Total}} = P_{\text{vap}} + P_{\text{air}}$$

Since the temperature of the air inside the flask is changing, the  $P_{\text{air}}$  needs to be calculated at each of the temperatures.

2. Use the gas law relationship shown below to calculate the  $P_{\text{air}}$  for trials 2-5. Remember that all gas law calculations require Kelvin temperature. Use the  $IP_{\text{air}}$  and  $IT_{\text{air}}$  from the data table to calculate the  $P_{\text{air}}$  for each trial using the corresponding Kelvin temperature for each trial.

$$\frac{IP_{\text{air}}}{IT_{\text{air}}} = \frac{P_{\text{air}(trial1,2,3,etc...)}}{T_{(trial1,2,3,etc...)}}$$

3. Calculate and record the  $P_{\text{vap}}$  for each trial:  $P_{\text{vap}} = P_{\text{Total}} - P_{\text{air}}$
4. Calculate and record  $\ln(P_{\text{vap}})$  and  $1/T$  (K) in the table below.

## DAY 1

	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
$P_{\text{total}}$ (kPa) (measured)					
$P_{\text{air}}$ (kPa) (calculated)					
$P_{\text{vap}}$ (kPa) (calculated)					
Temperature ( $^{\circ}\text{C}$ ) (measured)					
Temperature (K)					
$\ln(P_{\text{vap}})$					
$1/T$ (K)					

Sample Calculations for  $P_{\text{air}}$  and  $P_{\text{vap}}$



5. In order to determine the heat of vaporization,  $\Delta H_{\text{vap}}$ , a plot of  $\ln(P_{\text{vap}})$  vs. the reciprocal of absolute temperature ( $1/T$  in K) is needed.
6. Use an Excel spreadsheet to obtain the plot of  $\ln(P_{\text{vap}})$  vs.  $1/T$ .

- a. Enter the data ( $1/T$  and  $\ln(P_{\text{vap}})$ ) in columns A and B. Use Row 1 for column headings.

	A	B
1	$1/T$	$\ln(P_{\text{vap}})$
2		
3		
4		
5		
6		

- b. To draw a graph, select the x and y data, which would be data in fields A:2-6 and B:2-6.
- c. Click "Insert" and then "Chart". Choose "XY" scatter and select "Marked Scatter"
- d. When the graph is displayed, click on any data point on the chart and from the toolbar, select "Chart" and then "Insert Trendline".
- e. From the pop-up box, select the "Options" tab and check the boxes: 1) Display equation and 2) Display R-squared value and click OK.
- f. From the equation of the straight line, obtain the slope and set that equal to  $\frac{\Delta H}{R}$ . Using the fact that  $R = 8.314 \text{ J/molK}$ , calculate the value of  $\Delta H_{\text{vap}}$ .
7. The accepted value of the  $\Delta H_{\text{vap}}$  of ethanol is 42.32 kJ/mol. Calculate the percent error for your experimentally determined value of  $\Delta H_{\text{vap}}$  with the accepted value.
8. Repeat all the above calculations using the second set of data from Day 2

## DAY 2

	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
$P_{\text{total}}$ (kPa) (measured)					
$P_{\text{air}}$ (kPa) (calculated)					
$P_{\text{vap}}$ (kPa) (calculated)					
Temperature ( $^{\circ}\text{C}$ ) (measured)					
Temperature (K)					
$\ln(P_{\text{vap}})$					
$1/T$ (K)					

Sample Calculations for  $P_{\text{air}}$  and  $P_{\text{vap}}$