

## Experiment 12Z

WBH Edit 1/13/23

# INTERMOLECULAR FORCES AND THE LIQUID-VAPOR EQUILIBRIUM<sup>1</sup>

**MATERIALS:** tall 1000 mL beaker, 10 mL graduated cylinder with removable plastic stand, digital thermometer, iron tripod, wire screen, Bunsen burner, striker, graduated plastic syringes, 10 inch J-shaped injector needles/cannula, organic samples.

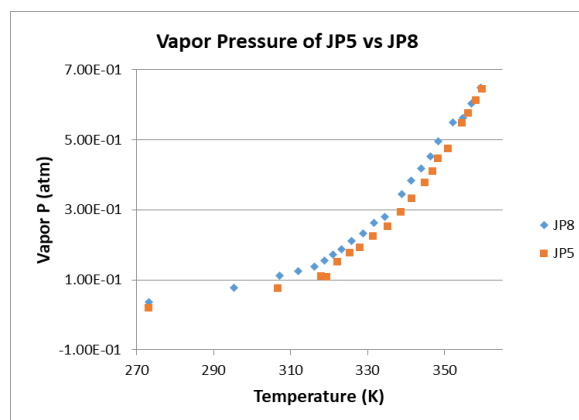
**PURPOSE:** The purpose of this experiment is to measure the effect of temperature on the vapor pressure of several liquids. The data will be analyzed to extract values for the heat of vaporization,  $\Delta_{\text{vap}}H$ , for each of the various liquids and these will be interpreted in terms of intermolecular forces.

**LEARNING OBJECTIVES:** By the end of this experiment, the student should be able to demonstrate the following proficiencies:

1. Use a spreadsheet program for data manipulation, graphing, and regression analysis.
2. Describe the effects of changes in temperature on the vapor pressure of a pure substance.
3. Describe how intermolecular forces influence the relative vapor pressure of a pure substance.
4. Understand the use of graphical methods to extract thermodynamic information from experimental pressure and temperature data.
5. Utilize Dalton's Law of Partial Pressures, and the Ideal Gas Law, to relate experimental data to properties of the test substance.

**DISCUSSION:** Any warship, but especially an aircraft carrier, carries the seeds of its own destruction deep within its hull; these seeds are the munitions and jet fuel used to carry out combat operations. See the following link for 5 minutes of gut-wrenching video of a fire aboard the USS Forrestal which killed 134 and injured 161 sailors over 50 years ago: <https://www.bing.com/videos/search?q=+uss+forrestal+mishap+video&view=detail&mid=81191609B9587C7B557D81191609B9587C7B557D&FORM=VIRE>

One way to minimize the danger of fuel fires aboard an aircraft carrier is the use of JP5, a less volatile jet fuel than JP8, the fuel used in commercial and land-based military aircraft. At the right are shown vapor pressure curves vs temperature for JP5 and JP8 determined by the procedure you will use today. To burn in an engine, any fuel must evaporate. As it evaporates, the resulting vapor exerts a pressure. Eventually the fuel vapor reaches a concentration (or pressure) where, when mixed with air, it can ignite and sustain combustion. An engine will not run if there is too little fuel (too "lean" a mixture) OR too much fuel (the engine is "flooded"). Note the exponential rise in the vapor pressures of both JP5 and JP8 but the slower rise in JP5. While the difference is small, the effect is significant – this is why JP5 is a less flammable (and safer) fuel.



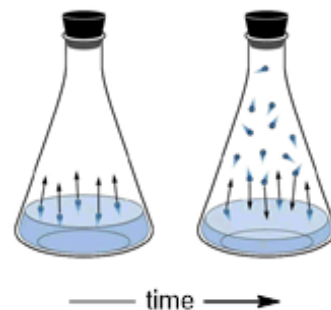
This raises two questions: 1) why does the liquid JP5 require a higher temperature to become a gas? and 2) why is the vapor P vs T graph curved (i.e. doesn't the Ideal Gas Law say P is directly proportional to T)? The answers lie in understanding the process of evaporation and the definition of vapor pressure.

The molecules of a gas move freely through the volume of their container, the individual molecules staying widely separated and experiencing little or no interaction with other molecules. Molecules in a liquid, while free to move throughout the volume of the sample, are constrained by intermolecular attractive forces to remain in contact with their neighbors. The strength of such intermolecular forces and the energy of motion available to the sample (based on the

<sup>1</sup>Fitzgerald, J.P., et. al. *J. Chem. Educ.* **2020**, *97*, 1406-1410; Levinson, G.S., *J. Chem. Educ.*, **1982**, *59*, 337.

temperature), together dictate the physical state of a substance. Evaporation, the conversion of a liquid to a gas, is an endothermic process since energy is required to overcome the attraction that a liquid molecule feels for its neighbors. The molar enthalpy of vaporization,  $\Delta_{\text{vap}}H$ , is the energy required to evaporate one mole of a substance at constant temperature and pressure. The magnitude of  $\Delta_{\text{vap}}H$  is thus a measure of the strength of the intermolecular forces in a pure substance. JP5 is a harder to evaporate because, in the liquid state, its molecules “feel” greater attraction for their neighbors than do the molecules of JP8 – thus a higher temperature is needed for molecules of JP5 to “break free” from their neighbors and “escape” into the gas phase. The molar enthalpy of vaporization,  $\Delta_{\text{vap}}H$ , for JP5 is larger than that of JP8.

The diagram illustrates what happens to a liquid in a closed container. The molecules in the liquid will have a distribution of energies at any temperature, Some of the molecules of the liquid (those in the higher energy range) will have sufficient energy to escape into the gas phase. Thus the pressure in the container will rise. Some of the gas phase molecules will hit the liquid surface and be unable to escape the attractions for their new neighbors; these (lower energy molecules) have undergone condensation and become part of the liquid. As more molecules accumulate in the gas phase (via evaporation), the rate of condensation will also increase. Eventually, the rate of evaporation and the rate of condensation will become equal, and the pressure in the container will level off at a constant value. The system is said to be in equilibrium, and the pressure of gas that exists over the liquid is called the **equilibrium vapor pressure** of the liquid. See the link below for an animation of this process:



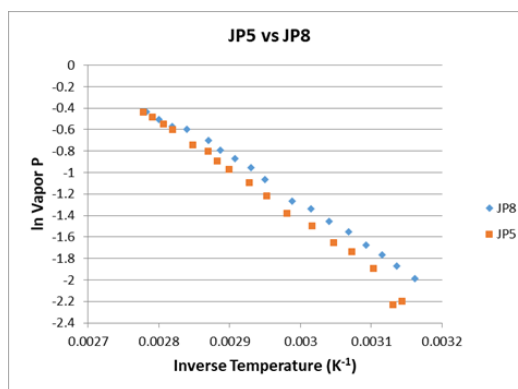
<https://www.youtube.com/watch?v=JsoawKguU6A>

Note that the same liquid has a different equilibrium vapor pressure at every temperature and, as shown in the JP5/JP8 diagram, vapor pressures increase exponentially with temperature. As the temperature increases, more molecules of liquid have enough energy to escape into the gas phase and so  $n_{\text{gas}}$  increases. Since two factors are increasing ( $T$  and  $n_{\text{gas}}$ ), the vapor pressure rises exponentially as predicted by the Ideal Gas Law,  $P = n_{\text{gas}} \cdot R \cdot T / V$ .

The equilibrium vapor pressure depends not only on the temperature of the sample, but also on the strength of the intermolecular forces holding molecules of the liquid together. For any individual substance, the variation of vapor pressure ( $P$ ) with temperature allows determination of the enthalpy of vaporization of that substance, as given by the Clausius-Clapeyron equation (assuming a constant  $\Delta_{\text{vap}}H$ ):

$$\ln P = \left( \frac{-\Delta_{\text{vap}}H}{R} \right) \cdot \left( \frac{1}{T} \right) + C \quad (1)$$

where  $R$  is the gas constant ( $8.314 \text{ J/mole}\cdot\text{K}$ ),  $T$  is the absolute temperature (in  $\text{K}$ ), and  $C$  is a constant. As seen in this equation, liquids with a large positive value of  $\Delta_{\text{vap}}H$  will have a low equilibrium vapor pressure at any temperature. As the temperature increases,  $\ln P$ , and thus  $P$ , also increases. Note that, based on Equation 1, a plot of  $\ln P$  vs  $1/T$  will be linear with a slope equal to  $-\Delta_{\text{vap}}H/R$ . Such a graph is shown to the right for JP5 and JP8. JP5, which has a lower vapor  $P$ , also has the steeper slope and therefore a higher  $\Delta_{\text{vap}}H$ . JP5 and JP8 are both mixtures of hydrocarbons, but they behave as pure substances according to Equation 1.



In this experiment, you will measure the equilibrium vapor pressure at several different temperatures for your assigned liquid. The liquids differ in the nature and strength of their intermolecular attractive forces. A small amount of your liquid (~1 mL) and a known amount of air will be trapped in an inverted graduated cylinder immersed in a beaker of water. Some of the liquid will evaporate until the equilibrium vapor pressure is established. The volume of the gas bubble (mixture of air and sample vapor) will be measured at several temperatures (always at atmospheric pressure). Since the moles of air in the trapped bubble is constant, the partial pressure of air in the mixture can be determined at any temperature. From this and the atmospheric pressure (measured with a barometer), the partial pressure (equal to the vapor pressure) of the sample at any temperature can be determined. Then the application of Eq. (1) allows the determination of the molar heat of vaporization  $\Delta_{\text{vap}}H$  by graphical methods.

**PROCEDURE:** (Work in pairs.)

1. Obtain the room air temperature and the barometric pressure and record it in the attached data sheet.
2. Put ~ 900 mLs of distilled water in a 1000 mL tall beaker. Remove the plastic bottom from a 10 mL graduated cylinder and fill it to the brim with water. Carefully cover the top of the cylinder with your finger, invert it, and place it in the beaker without allowing any air into the cylinder. The cylinder must be immersed in the water bath throughout the experiment.
3. Using a plastic syringe and a J-shaped needle/cannula, inject 2.00 mLs of air (measured accurately!) into the cylinder. Do this by positioning the tip of the needle/cannula inside the graduated cylinder before depressing the syringe plunger. When the syringe plunger is depressed, the injected air will rise through the water and be trapped in the graduated cylinder.
4. *ORGANIC STEP:* See your Instructor for a sample loading syringe containing your assigned organic liquid. (If assigned water as a sample, just continue to step 5.) You or your instructor should load the sample by carefully placing the opening of the needle/cannula tip inside the submerged cylinder. Slowly inject ~ 1 mL of organic liquid; it will rise through the water in the cylinder and form a separate layer under the air bubble.
5. Record the identity of your sample liquid and its normal boiling point on the data sheet. (See the table on p. E12Z-6. You should have this value in mind as you work through the procedure.) Assemble the apparatus as demonstrated by your instructor. Place the wire screen on top of the iron tripod and place the 1 L beaker on the wire screen. Place a ring stand with clamp near the tripod and use it to suspend your thermometer in the beaker, with the tip about mid-height. Finally, set up a Bunsen burner under the apparatus and begin heating.
6. Bend the wide blade of your metal spatula to a shallow angle (roughly 30°). Using the bent spatula, **continuously stir the water while it heats**. This is done to ensure that the temperature of the water bath is uniform throughout. The bend will help lift and distribute the hotter water from the bottom.
7. As the apparatus begins heating, fill a 250 mL beaker with crushed ice.
8. While continuing to stir, monitor the volume of the gas bubble within the submerged cylinder. As the temperature rises, the gas bubble will expand exponentially – it must not push any of the organic liquid out of the cylinder. (**Should this happen, immediately shut off the gas to the Bunsen burner and notify your instructor**). Continue to stir and heat until the gas bubble passes the 9 mL mark on the cylinder. Typically, you need to heat the water bath to a temperature ~10°C below the boiling point of your organic liquid, but focus on the *size of the bubble*, not this temperature guideline.
9. As the volume of the gas bubble *passes* the 9 mL mark while expanding, shut off the gas to the Bunsen burner. Continue to stir. The bubble will continue to expand slightly after the heating is stopped as the vapor pressure reaches its equilibrium value. If the bubble volume get so large that it threatens to push some of the organic liquid out of the cylinder, you may add a small “pinch” of ice to speed the cooling process.
10. As the water bath cools, the gas bubble will begin to contract. Once the gas bubble begins contracting, begin to record the bath temperature (to the nearest 0.1°C and bubble volume (to the nearest (0.05 mL). Continue to stir throughout the measurement process. Collect 10-12 temperature-volume measurements for bubble volumes between 10 and 3 mLs. Note that most of the volume change occurs in the higher temperature range. Small “pinches” of ice may be added between measurements to speed the cooling process.
11. If time allows, collect additional temperature-volume measurements for bubble volumes less than 3 mLs. Do this by a) adding ice to the water bath and/or b) pouring off some of the bath water. When pouring off bath water, do not introduce more air into the graduated cylinder or allow the organic liquid to escape from the cylinder. Collecting data over a larger temperature range better illustrates the exponential shape of the vapor pressure vs temperature graph.
12. The organic liquid must be collected for disposal at the end of the experiment. Quickly lift the inverted graduated cylinder out of the apparatus and flip it over. The organic liquid will be contained in the graduated cylinder and must be poured into a waste collection container. The remaining water in the 1 L beaker is not contaminated with organic and can be poured down the drain. The graduated cylinder should be inverted in a test-tube rack and allowed to dry before the next class.

Name \_\_\_\_\_

Section \_\_\_\_\_

Lab Partner \_\_\_\_\_

Date \_\_\_\_\_

**DATA SECTION**  
**Experiment 12Z**

Barometric P (mm Hg) \_\_\_\_\_

Room Air T (°C) \_\_\_\_\_

Sample Liquid \_\_\_\_\_

Sample Boiling T (°C) \_\_\_\_\_

Temperature (°C)	Gas Bubble Volume (mL)

## DATA TREATMENT

### Experiment 12Z

1. Set up an Excel spreadsheet with the experiment title and your name at the top. In separate rows, enter labels for the barometric pressure, room air temperature and organic liquid (if any) and enter your values for each quantity in its own cell. Be sure to include the units of each measured value.
2. Using a formula in your spreadsheet, determine the number of moles of air,  $n_{\text{air}}$ , trapped in the gas bubble. Treat the gas bubble as an ideal gas:  $n_{\text{air}} = P_{\text{bar}}V/RT$ , where the temperature value is that of the air in the room and the pressure is the barometric pressure  $P_{\text{bar}}$ . Include a label for this value with units in your spreadsheet.

Show your calculation of  $n_{\text{air}}$ .

$$n_{\text{air}} = \text{_____ mol air}$$

3. In your spreadsheet, create columns for the Volume (mL) and Temperature ( $^{\circ}\text{C}$ ) data you collected and enter your volume and temperature data in the appropriate columns. A small correction must be made to the measured gas volumes due to the liquid meniscus. Label another column in your spreadsheet as Corrected Volume (mL). Use an Excel formula to subtract 0.2 mLs from your measured volumes and enter these values in this column.
4. Create additional columns in your spreadsheet for Corrected Volume (in L) and Absolute Temperature (in K). Use Excel formulas to convert all data from previous columns and enter the data in the new columns.
5. At all temperatures, the gas bubble is a mixture of air *plus* sample vapor. The partial pressure of the air in the bubble will change because the temperature and volume change, even though the amount of air (the number of moles of air) is constant. Add a column  $P_{\text{air}}$  to your spreadsheet and for each of the data rows in the spreadsheet, set up an Excel formula to calculate the partial pressure of air in the gas bubble:  $P_{\text{air}} = n_{\text{air}}RT/V$ , recognizing that the number of moles of air does not change from step 2 above. Note that  $P_{\text{air}}$  should be calculated in units of atmospheres.

Show a sample calculation of  $P_{\text{air}}$  for your highest temperature data point.

$$P_{\text{air}} = \text{_____ atm air}$$

6. By Dalton's Law, the total pressure of a gas mixture is the sum of the partial pressures of the components. Your mixture contains air plus the sample vapor, and the total pressure of the mixture is the measured barometric pressure. Add a column  $P_{\text{sample}}$  to your spreadsheet and for each of the data rows, use Dalton's Law to calculate the partial pressure of the test sample,  $P_{\text{sample}}$  in the gas bubble:  $P_{\text{sample}} = P_{\text{bar}} - P_{\text{air}}$ . Note that  $P_{\text{sample}}$  should be calculated in units of atmospheres. Show a sample calculation for your highest temperature data point.

$$P_{\text{sample}} = \text{_____ atm}$$

7. Add columns for the natural log of the sample pressure,  $\ln(P_{\text{sample}})$ , and the inverse Kelvin temperature ( $1/T(\text{K}^{-1})$ ) to your spreadsheet. Use Excel formulas to calculate these quantities from the other entries.

8. Plot the vapor pressure curve  $P_{\text{sample}}$  vs.  $T(\text{K})$  for your test substance. Adjust the x and y axes limits so that the experimental data fills the graph. Does  $P_{\text{sample}}$  increase exponentially with  $T$ ? Do not be alarmed if your data looks linear – you need to plot vapor pressure over a temperature range of  $40^\circ \text{K}$  or more to see the exponential curve.
9. Select **only** the points where the gas bubble was 3.0 mL or greater. Plot  $\ln(P_{\text{sample}})$  vs.  $1/T (\text{K}^{-1})$  for these points. Again adjust the x and y axes limits so that the experimental data fills the graph. Perform a **linear** trendline analysis on this plot, showing the trendline equation and  $R^2$  value on the plot. Enter your trendline equation and  $R^2$  values here. (The reason for limiting the selection is that the error in  $P_{\text{sample}}$  increases as the bubble volume decreases / approaches 2.0 mL.)

Equation of trendline: \_\_\_\_\_ Slope (with units): \_\_\_\_\_

$R^2$  for trendline: \_\_\_\_\_

10. According to Equation (1), the slope of your plot is related to the heat of vaporization of the sample. From the slope of your regression line, determine the molar enthalpy (heat) of vaporization,  $\Delta H_{\text{vap}}$ , and compare it to the accepted value of your test sample (see table below). Calculate the percent error. Show your work.

sample \_\_\_\_\_ measured  $\Delta H_{\text{vap}}$  \_\_\_\_\_ kJ/mol % error \_\_\_\_\_

11. The “normal boiling point” of a substance is the temperature at which the vapor pressure of the substance equals 1 atm. Use the trendline equation just determined to find the normal boiling point (i.e. the  $T$  where the  $P = 1$ ) of your sample. Calculate the % error from the literature value (see table). Show your work.

sample \_\_\_\_\_ measured normal boiling point \_\_\_\_\_ K % error \_\_\_\_\_

12. **Enter your result in the Class Data table.** Record data for the other samples from the Class Data table, so that you have experimental  $\Delta_{\text{vap}}H$  values for all samples used.

Literature Values:

Sample	Formula	Molar Mass (g/mol)	Normal Boiling Point <sup>1</sup> (°K)	Heat of Vaporization <sup>2</sup> $\Delta H_{\text{vap}}$ (kJ/mol)
Water	H <sub>2</sub> O	18.02	373.2	40.7
Methyl t-Butyl Ether	C <sub>5</sub> H <sub>12</sub> O	88.15	328.2	30.1
n-Hexane	C <sub>6</sub> H <sub>14</sub>	86.18	341.9	31.6
Ethyl Acetate	C <sub>4</sub> H <sub>8</sub> O <sub>2</sub>	88.11	350.2	35.4
n-Heptane	C <sub>7</sub> H <sub>16</sub>	100.20	371.5	36.1
n-Octane	C <sub>8</sub> H <sub>18</sub>	114.23	398.7	41.6
2-Pentanol	C <sub>5</sub> H <sub>12</sub> O	88.15	392.0	51.9

<sup>1</sup>Glushko ThermoCenter, “Entropy and Heat Capacity of Organic Compounds” in **NIST Chemistry WebBook, NIST Standard Reference Database Number 69**, Eds. P.J. Linstrom and W.G. Mallard, National Institute of Standards and Technology, Gaithersburg MD, 20899, <http://webbook.nist.gov>, (retrieved April 8, 2016).

<sup>2</sup>Chickos, J.S. and Acree Jr., W.E, *J. Phys. Chem. Ref. Data*, 32, 519 (2003), averages from data sets over temperature range ~280K-350K

CLASS DATA  
Experiment 12Z

Sample	Heat of Vaporization ( $\Delta H_{\text{vap}}$ ) (kJ/mol)				Average (kJ/mol)
water					

**QUESTIONS**  
**Experiment 12Z**

1. The table shows the structures of some substances which may be used in this experiment

Name	Chemical Formula	Condensed Formula	Structural Formula
Water	H <sub>2</sub> O	H <sub>2</sub> O	
Ethyl acetate	C <sub>4</sub> H <sub>8</sub> O <sub>2</sub>	CH <sub>3</sub> COOCH <sub>2</sub> CH <sub>3</sub>	
n-Hexane	C <sub>6</sub> H <sub>14</sub>	CH <sub>3</sub> (CH <sub>2</sub> ) <sub>4</sub> CH <sub>3</sub>	
Methyl t-butyl ether	C <sub>5</sub> H <sub>12</sub> O	(CH <sub>3</sub> ) <sub>3</sub> COCH <sub>3</sub>	
n-Octane	C <sub>8</sub> H <sub>18</sub>	CH <sub>3</sub> (CH <sub>2</sub> ) <sub>6</sub> CH <sub>3</sub>	
2-Pentanol	C <sub>5</sub> H <sub>12</sub> O	CH <sub>3</sub> CH(OH)(CH <sub>2</sub> ) <sub>2</sub> CH <sub>3</sub>	
n-Heptane	C <sub>7</sub> H <sub>16</sub>	CH <sub>3</sub> (CH <sub>2</sub> ) <sub>5</sub> CH <sub>3</sub>	

Using the structures provided, circle **all** of the intermolecular forces present in each pure liquid.

Water -	Dipole-Dipole	Hydrogen Bonding	London Dispersion Forces
Ethyl acetate -	Dipole-Dipole	Hydrogen Bonding	London Dispersion Forces
n-Hexane -	Dipole-Dipole	Hydrogen Bonding	London Dispersion Forces
Methyl t-butyl ether -	Dipole-Dipole	Hydrogen Bonding	London Dispersion Forces
n-Octane -	Dipole-Dipole	Hydrogen Bonding	London Dispersion Forces
2-Pentanol -	Dipole-Dipole	Hydrogen Bonding	London Dispersion Forces
n-Heptane -	Dipole-Dipole	Hydrogen Bonding	London Dispersion Forces

2. For a clear comparison, focus on the substances of question 1 that can utilize **ONLY** London dispersion forces. What trend would be expected for the normal boiling points and heats of vaporization of these substances? Is that trend observed in the literature (not experimental) data? (see table on p. E12Z-6). **EXPLAIN YOUR ANSWER.**



3. For that same set of substances that utilize ONLY London dispersion forces, what trend would you expect for the value of the equilibrium vapor pressure of the substance at 25°C? EXPLAIN YOUR ANSWER.

4. Consider the liquids actually used in your experiment, and tabulated in the Class Data table. Use the averages of the **measured** heats of vaporization, and rank the substances from greatest to least  $\Delta H_{\text{vap}}$ .

highest  $\Delta H_{\text{vap}}$  \_\_\_\_\_ lowest  $\Delta H_{\text{vap}}$

5. Consider the substances actually measured, and the experimental rankings of question 4. How well do the measured values match your expectation based on the intermolecular forces you identified for the substances? EXPLAIN YOUR ANSWER.

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Name \_\_\_\_\_

Section \_\_\_\_\_

Date \_\_\_\_\_

**PRE-LAB Questions**  
**Experiment 12Z**

Use the following table to consider the effect of structure on intermolecular forces.

methyl t-butyl ether	pentane	pentanoic acid	pentanol
C <sub>5</sub> H <sub>12</sub> O	C <sub>5</sub> H <sub>12</sub>	C <sub>5</sub> H <sub>10</sub> O <sub>2</sub>	C <sub>5</sub> H <sub>12</sub> O

1. Which of the following pure liquids would be expected to utilize ONLY London Dispersion Forces? Circle ALL that apply.)

- a) methyl t-butyl ether (C<sub>5</sub>H<sub>12</sub>O)      b) pentane (C<sub>5</sub>H<sub>12</sub>)      c) pentanoic acid (C<sub>5</sub>H<sub>10</sub>O<sub>2</sub>)      d) pentanol (C<sub>5</sub>H<sub>12</sub>O)

2. Which of the following pure liquids would be expected to utilize Dipole-Dipole forces, but NOT Hydrogen Bonding? (Circle ALL that apply.)

- a) methyl t-butyl ether (C<sub>5</sub>H<sub>12</sub>O)      b) pentane (C<sub>5</sub>H<sub>12</sub>)      c) pentanoic acid (C<sub>5</sub>H<sub>10</sub>O<sub>2</sub>)      d) pentanol (C<sub>5</sub>H<sub>12</sub>O)

3. Which of the following pure liquids would be expected to utilize Hydrogen Bonding? (Circle ALL that apply.)

- a) methyl t-butyl ether (C<sub>5</sub>H<sub>12</sub>O)      b) pentane (C<sub>5</sub>H<sub>12</sub>)      c) pentanoic acid (C<sub>5</sub>H<sub>10</sub>O<sub>2</sub>)      d) pentanol (C<sub>5</sub>H<sub>12</sub>O)

4. Which of the following alkanes will have the highest normal boiling point?

- a) C<sub>2</sub>H<sub>6</sub>                      b) C<sub>4</sub>H<sub>10</sub>                      c) C<sub>6</sub>H<sub>14</sub>                      d) C<sub>8</sub>H<sub>18</sub>                      e) C<sub>10</sub>H<sub>22</sub>

5. Which of the following alkanes will have the highest vapor pressure at room temperature?

- a) C<sub>2</sub>H<sub>6</sub>                      b) C<sub>4</sub>H<sub>10</sub>                      c) C<sub>6</sub>H<sub>14</sub>                      d) C<sub>8</sub>H<sub>18</sub>                      e) C<sub>10</sub>H<sub>22</sub>

6. For a substance studied in Exp. 12Z, a plot of  $\ln P$  vs  $1/T$  (in K) had the trendline equation  $y = -4525.9x + 12.916$ . Based on Equation 1 in the lab, use the slope of the line to determine the value of  $\Delta H_{\text{vap}}$  for the substance, in kJ/mol.

7. The “normal boiling point” of a substance is the temperature at which the vapor pressure of the substance equals 1 atm. Based on the trendline equation data in question 6, what is the temperature (in K) of the normal boiling point?