Vapor Pressure of Liquids Lab
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**Abstract**

 In this experiment, the purpose was to investigate the relationship between the vapor pressure of a liquid and its temperature. The purpose of the extension was to prove this relationship using the Clausius-Clapeyron equation, including data found during the main experiment procedure(s).

 The experimental results obtained supported our hypothesis in this procedure the relationship of vapor pressure to temperature was found. The hypothesis was that the two bore a direct relationship to each other, which means one increasing as the other increases. The data collected supports the original hypothesis. As vapor pressure went up, so did the temperature. One can also infer that increasing the temperature and thus, kinetic energy, increases the vapor pressure and the rate at which random interactions occur. Using this information, one would predict that raising the temperature in a bottle of soda causes the vapor pressure to increase and more molecules to escape faster, leading the soda to go “flat” at a much faster rate. Based on the results, methanol and ethanol exhibit different kinds of intermolecular forces and different vapor pressures. The heat of vaporization (∆H*vap*) was found to be 162, which based off of the established value of 37.43, constituted a 332.81% error.

**Introduction**

Vapor pressures are an everyday surrounding to us, but many of us do not realize this common occurrence. The simplest way to define vapor pressure is the pressure created by a *vapor* or gas that forms above a liquid or solid of the same substance.1 Picturing a soda bottle, the bottle itself is pressurized until the hiss sound you hear upon opening the bottle for the very first time. That is the vapor escaping. In a closed soda bottle, gas molecules are converting to liquid and liquid molecules are converting to gas in dynamic equilibrium. Temperature changes (such as placing the drink in a refrigerator) are used to make sure the soda does not go “flat” (which can be achieved simply by leaving the bottle open in a room). The vapor pressure of the soda is the pressure exerted above the liquid portion. Closely related to vapor pressure, is *atmospheric pressure* which is the pressure exerted in an open environment at room temperature. The atmospheric pressure is often used to calculate the vapor pressure for a particular experiment substance. Relevance and application of vapor pressure can also vary, depending on what you are using it for. For instance, one of the most known things in science is that a higher temperature possesses higher kinetic energy, or energy of motion. Well, vapors are *gases*, that is they exhibit a high kinetic energy, generally colliding with the sides of a container (if closed container) and dispersing freely into the environment (if open container).

 The importance of vapor pressure simply relies on the amount of pressure being applied. In boiling points, the amount of vapor pressure directly influences the boiling point. Commonly the example of vapor pressure is that the boiling point of water in Denver, CO is lower than one at sea level due to decreased atmospheric pressure.2 The United States Department of Agriculture (USDA) even has an online directory of the various vapor pressures, proving the widespread use of vapor pressure. Things get a little more complicated when dealing with vapor pressures of mixtures. This is where Raoult’s Law comes into play. Vapor pressure relies on the existence of three distinct types of intermolecular forces – London forces, present in all molecules, dipole-dipole interactions, which are the result of a polarized structure, and hydrogen bonds, which are the result of a hydrogen atom covalently bonded to a highly electronegative atom (such as oxygen) in a polarized structure. The more intermolecular forces a molecule has, the lower the vapor pressure of the molecule and the lower the boiling point of the substance. The boiling point is especially critical in commercial food applications, fast food franchising, etc. The use of vapor pressure far exceeds the food industry and even includes the automobile manufacturing industry, consumer goods manufacturing, consumable drinks manufacturing, chemistry lab equipment and lab equipment manufacturing, and even the use of some general carpenter tools, such as levels (how do you think they “level” a level to begin with?). General rule of thumb is if the substance is high in vapor pressure, then the solution is volatile.

**Experimental**

The experiment was taken from the LabQuest 10: Vapor Pressure of Liquids Lab experiment handout from Vernier Software.3

**Data / Calculations**

In this experiment, you must account for the variation of pressure inside the flash, regardless of methane presence, due to the higher temperature(s) of the water. This is where the gas law equation is helpful. To use the gas law equation, you must first convert all of your temperatures from degrees Celsius (° C) to degrees Kelvin (° K).

Now using the gas law equation, substitute the *P1* with the value of measured atmospheric pressure and *T1* with the room temperature value. You must then plug in the temperature value for each water bath and solve for *P2.*

After finding the adjusted pressure value, you now subtract the adjusted pressure value from the measured pressure value to find the vapor pressure.

 In the extension to the main experiment, the Clausius-Clapeyron equation is used to describe the relationship between the vapor pressure and the absolute temperature.

The equation is translated into slope-intercept form (y = mx + b) to make the graph easier to read and interpret.

After finding the heat of vaporization, compare the result to the established value (37.43) 4 and find your percent error.

**Conclusion**

In conclusion, the vapor pressures of methanol and ethanol were used to describe the relationship between vapor pressure and absolute temperature, which was found to be a direct relationship. The experiment also showed that methanol and ethanol have different intermolecular forces at work and different vapor pressures. The amount of intermolecular forces and the force(s) they exhibit affects the vapor pressure. Remembering from the introduction, the three forces that can affect the vapor pressure are London forces (which are exhibited in all), dipole-dipole forces (which exist in polar substances), and hydrogen bonds (which require a polar structure, and a hydrogen atom covalently bonded to a highly electronegative atom such as fluorine, chlorine, or oxygen). All in all, the relationship was nonetheless discovered and described, supporting the original hypothesis, and showed just how widespread the use of vapor pressure really is.

**Graphs and Figures**

**\* The fourth data point is not graphed because it lies outside the logical data points.**

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| --- | --- | --- | --- |
| Temperature ( °C ) | Vapor Pressure ( kPa ) | 1 / Kelvin ( 1 / °K ) | ln Pressure |
| 6.7 | -6 | 4 x 10-3 | 1.90 |
| 17.5 | -3 | 3 x 10-3 | 2.86 |
| 20.7 | 2.22 | 3 x 10-3 | 3.03 |
| 35.3 | 14 | 3 x 10-3 | 3.56 |

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**References**

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