

Enthalpy of vaporization of water analysis engineering essay



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The purpose of this experiment is to determine the enthalpy of vaporization of water by taking note of the vapor pressure of water over many temperatures and applying the Clausius-Clapeyron equation.

Data

Temperature - Volume Observations

Temperature (*C)

Volume of Air (ml)

78.0

10.01

76.0

9.41

74.0

8.97

72.0

8.43

70.0

8.03

68.0

7. 68

66. 0

7. 38

64. 0

7. 19

62. 0

6. 94

60. 0

6. 71

58. 0

6. 52

56. 0

6. 38

54. 0

6. 09

52. 0

6. 05

50. 0

5. 99

5. 0

4. 62

4. 9

4. 60

Initial Conditions

Initial Volume of Air:

4. 8 ml

Larger Beaker Water Volume:

~800 ml

Initial Temperature:

23. 0 °C

Atmospheric Pressure:

74. 44 cm Hg

ΔH_{vap} for water

40. 65 kJ/mol

Water was cooled with ice for about 20 minutes. Two measurements were made to ensure that the temperature and air volume had stabilized.

Graphs and Table of Results

Table of Calculated Results

Reg. Eq: $y = -5790x + 15.7$

Slope: -5790

Std. Dev (Slope): 152

$n_{air} = 1.9 \times 10^{-4} \text{ mol}$

Calculated = 49.6 kJ/mol

Calculated Std. Dev: 1.26 kJ/mol

Theoretical = 40.65 kJ/mol

Percent Deviation = 22.0%

Calculations

Temperature converted from degrees Celsius to degrees Kelvin.

$78.0^\circ\text{C} + 275.15 = 351.2^\circ\text{C}$

To correct for the reversed meniscus, volume readings are reduced by 0.2 ml.

$10.01 \text{ ml} - 0.2 = 9.8 \text{ ml}$

Atmospheric Pressure is converted from cm Hg to atmospheres

To determine the number of moles of trapped air, the ideal gas equation was used with the volume reading from the lowest temperature (5°C).

$$PV = nRT$$

$$P = 0.9795 \text{ atm}, V = , R = 0.0821 \text{ L atm K}^{-1}\text{mol}^{-1} T = 278.2 \text{ K}$$

$$0.9795 \times 0.0044 = n \times 0.0821 \times 278.2$$

$$n = 1.9 \times 10^{-4} \text{ mols air}$$

The partial pressure of air at any temperatures is determined by using the ideal gas equation.

$$PV = nRT$$

$$V = 0.0098 \text{ l}, n = 1.9 \times 10^{-4} \text{ mols air}, R = 0.0821 \text{ L atm K}^{-1}\text{mol}^{-1}, T = 351.2 \text{ K}$$

$$P \times 0.0098 = (1.9 \times 10^{-4}) \times 0.0821 \times 351.2$$

$$P = 0.56 \text{ atm}$$

Partial pressure of water at any temperature is determined by subtracting the partial pressure of air from the atmospheric pressure.

$p_{\text{air}} + p_{\text{water}} = P_{\text{atm}}$, p_{air} - partial pressure of air, p_{water} - partial pressure of water, P_{atm} - atmospheric pressure

$$p_{\text{air}} = 0.56, p_{\text{atm}} = 0.98$$

$$p_{\text{water}} = 0.98 - 0.56 = 0.42 \text{ atm}$$

partial pressure of water is the vapor pressure at a set temperature: $p_{\text{water}} = p_{\text{vap}}$

, so when plotting $\ln(p_{\text{vap}})$ vs. $1/T$, slope will be

Slope is $-5790 =$

$$= -(-5790) \cdot 8.314 \text{ J/molK} = 49600 \text{ J/mol} = 49.6 \text{ kJ/mol}$$

Used LINEST function. Std. Dev of slope = 152 = -std. dev enthalpy/R

$$152 = -\text{std. dev enthalpy} / 8.314$$

$$\text{Std. dev} = 1260 \text{ J/mol} = 1.26 \text{ kJ/mol}$$

Percent Deviation

$$\text{Actual enthalpy} = 49.6 \text{ kJ/mol}, \text{ Theoretical enthalpy} = 40.65 \text{ kJ/mol}$$

Discussion

For a substance to change physical state from liquid to gas, there must be an addition of heat to the substance. The added heat will agitate the molecules and break intermolecular attractions, allowing the molecules to become a gas. In this experiment, liquid water was heated to a gas: $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$. The amount of energy per mole required to vaporize water is known as the enthalpy of vaporization of water, and that is what was calculated in this

experiment. The enthalpy of vaporization, temperature, and vapor pressure are related by the Clausius-Clapeyron equation: is thus the slope of the linear plot of $\ln(p_{\text{vap}})$ and $1/T$. To attain this linear plot, $\ln(p_{\text{vap}})$ needs to be calculated over a range of temperatures.

An inverted graduated cylinder placed in a larger beaker of water trapped a constant amount of air molecules. When the water was cooled to low temperatures, the amount of water vapor in the graduated cylinder was minimized. At these conditions, the ideal gas law was applied to determine the number of moles of air trapped in the graduated cylinder. The volume of the trapped air was also measured as temperature varied from 50°C to 80°C. With the knowledge of the number of moles of trapped air, the ideal gas law was utilized once more to determine the partial pressure of air over the temperature range of 50°C to 80°C. It is the combination of the air pressure and water vapor pressure in the graduated cylinder that resists the opposing atmospheric pressure. As such the difference between the atmospheric pressure and the air pressure yielded the vapor pressure over the range of temperatures. Taking the natural log of these vapor pressure and plotting them against the reciprocal of the temperature values yields a linear plot where the slope is, as described by the Clausius-Clapeyron equation. Hence is calculated by multiplying the slope value by $-R$, which is 8.314 J/(K·mol). was successfully calculated to be 49.6 kJ/mol by observing vapor pressures over the range of temperatures 50°C to 80°C and applying the Clausius-Clapeyron equation, fulfilling the purpose of this experiment.

Error and Uncertainty Analysis

The standard deviation of the calculated enthalpy of vaporization is 1.26 kJ/mol, which represents a 2.54% uncertainty in the enthalpy value. The percent deviation is 22.0% from the literature value for the enthalpy, which is 40.65 kJ/mol.

There were a number of instrumental uncertainties in this experiment. The digital thermometer used had an accuracy of $\pm 0.1^\circ\text{C}$, yielding an uncertainty of between 0.125% and 0.2% in the measured temperature ($0.1/50 \times 100$ and $0.1/80 \times 100$). The 10 ml graduated cylinder used had an accuracy of ± 0.1 ml, yielding an uncertainty of between 1.0% and 1.67% in the measured volume ($0.1/10.01 \times 100$ and $0.1/5.99 \times 100$). The barometer used to measure atmospheric pressure had an accuracy of ± 0.01 cm Hg, yielding an uncertainty of $\pm 0.013\%$ in the measured atmospheric pressure. When inverting the graduated cylinder, an inverse meniscus occurred, and this was corrected for by deducting 0.2 ml from all volume measurements.

There were no errors in the execution of the experiment, however, there were some assumptions made. Firstly, this experiment assumed that air is an ideal gas, which it is definitely not. Air is made up of a combination of nitrogen, oxygen, and argon among many other elements. To correctly determine the number of moles of air trapped in the graduated cylinder as well as more accurate values for the partial pressure of air, van der Waals equation should be used. The experiment was conducted with tap water, and not distilled water, and as such the actual enthalpy of vaporization might be higher than the enthalpy of vaporization of water (due to the presence of other compounds that increase the intermolecular attractions of the water).

Another assumption made in this experiment is that at approximately 5°C ,
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the graduated cylinder contained a negligible amount of water vapor. If water vapor were still present in the graduated cylinder, then the value of p_{air} would be lower than calculated, yielding a higher p_{vapor} , and a higher enthalpy of vaporization.

The most important source of error is the assumption that air is an ideal gas.

Air is a mixture of multiple elements, and is very far from ideal in nature.

Making such the assumption that air is ideal is not only a significant one to take, but also affects the manner of calculation of the number of moles of air in the graduated cylinder and the partial pressure of air, two values critical to determining the enthalpy of vaporization.