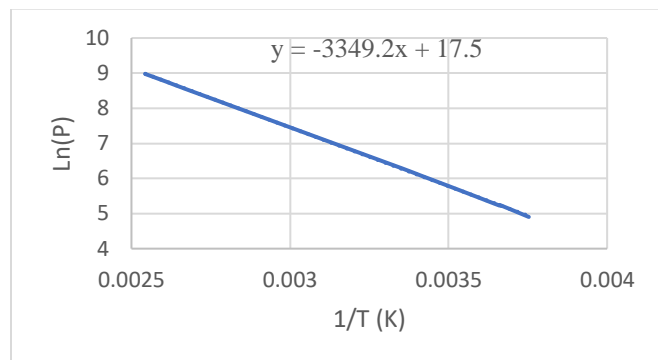


The Clausius-Clapeyron Equation: Enthalpy of Vaporization

1. Use the following plot of $\ln(P)$ vs $1/T$ to answer the questions. The pressures have units of mmHg.



a) What is the pressure at a temperature of 34.7°C ?

$$34.7^\circ\text{C} = 307.9 \text{ K} \quad \frac{1}{T} = \frac{1}{307.9 \text{ K}} = 0.003248 \text{ K}^{-1}$$

$$y = -3349.2 \times 0.003248 \text{ K}^{-1} + 17.5 = 6.62$$

$$\ln(P) = 6.62 \quad \text{and} \quad e^{6.62} = \mathbf{750 \text{ mmHg}}$$

b) What is the pressure, in mmHg, at the normal boiling point?

The pressure is 760 mmHg at the normal boiling point

c) What is the enthalpy of vaporization, ΔH_{vap} ?

$$\text{Slope} = -\frac{\Delta H_{\text{vap}}}{R}$$

$$3349.2 \times 8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}} = 27845 \frac{\text{J}}{\text{mol}} = \mathbf{27.8 \text{ kJ/mol}}$$

d) Use the Clausius-Clapeyron equation to determine the enthalpy of vaporization of diethyl ether under the following conditions: The vapor pressure is 135 mmHg at -6.7°C , and 438 mm Hg at 19.9°C .

$$\ln\left(\frac{P_2}{P_1}\right) = -\frac{\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1}\right) \quad P_2 = 438 \text{ mmHg} \quad T_2 = 293.05 \text{ K}$$

$$P_1 = 135 \text{ mmHg} \quad T_1 = 266.45 \text{ K}$$

$$\ln\left(\frac{438 \text{ mmHg}}{135 \text{ mmHg}}\right) = -\frac{\Delta H_{\text{vap}}}{8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}}} \left(\frac{1}{293.05 \text{ K}} - \frac{1}{266.45 \text{ K}}\right) \quad \text{Solve for } \Delta H_{\text{vap}}$$

$$\Delta H_{\text{vap}} = 28814 \text{ J/mol} = \mathbf{28.8 \text{ kJ/mol}}$$