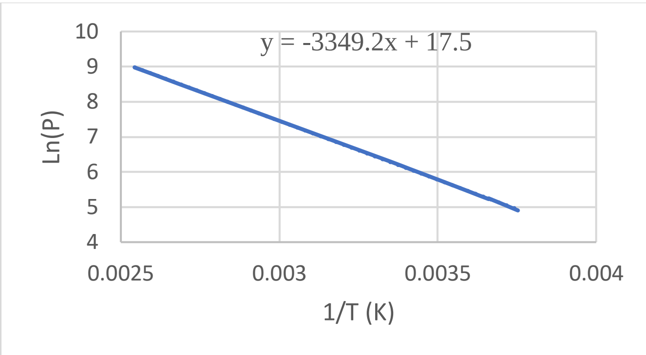
The Clausius-Clapeyron Equation: Enthalpy of Vaporization

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?

34.7oC = 307.9 K $\frac{1}{T} = \frac{1}{307.9\,K} = 0.003248\,K^{-1}$

y = -3349.2 x 0.003248 K^{-1} + 17.5 = 6.62

ln(P) = 6.62 and e^$^{6.62}$ = **750 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}?

Slope = $-\frac{\Delta Hvap}{R}$

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

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.

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

P_1 = 135 mmHg T_1 = 266.45 K

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

ΔH_{vap} = 28814 J/mol = **28.8 kJ/mol**