CH353 – Physical Chemistry I
Spring 2013, Unique 52575

Quiz 4, 1 March 2013

The molar enthalpy of vaporization of chloroform ($CHCl_3$) at its normal boiling point, 80.1°C, is 30.7 kJ mol^{-1}. Assuming that ΔH_{vap} and ΔS_{vap} remain constant, determine ΔG_{vap} at 75°C, 80.1°C, and 85°C.

$\Delta H_{vap}^\circ = 30.7\ kJ/mol$, $T_{vap} = 80.1°C = 353.1\ K$

$\Delta G_{vap} = \Delta H_{vap} - T_{vap}\Delta S_{vap}$

@ $T = 353.1\ K$, $\Delta G_{vap} = \Delta G_{vap}^\circ = 0$ (by definition, the transition is in equilibrium)

$\Delta G_{vap} = \Delta H_{vap}^\circ - T_{vap}\Delta S_{vap}^\circ$, $\Delta S_{vap}^\circ = \frac{\Delta H_{vap}^\circ}{T_{vap}} = \frac{30.7\times10^3\ J/mol}{353.1\ K}$

$\Delta S_{vap}^\circ = 86.9\ J/K\,mol$ assume this is constant

@ $T = 348.1\ K$,

$\Delta G_{vap} = (30.7\times10^3\ J/mol) - (348.1K)(86.9\ J/K\,mol) = 441\ J/mol$

@ $T = 358.1K$,

$\Delta G_{vap} = (30.7\times10^3\ J/mol) - (358.1K)(86.9\ J/K\,mol) = -428\ J/mol$