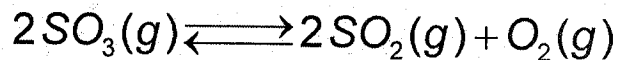


Chap 17 - Feb 14 - Slide 59

$$\ln(K_{c,2}/K_{c,1}) = -\frac{\Delta H^\circ}{R} \left[\frac{1}{T_2} - \frac{1}{T_1} \right] \quad R = 8.31 \text{ J/mol-K}$$

To illustrate, consider the equilibrium:



The equilibrium constant for this reaction is 3.6×10^{-3} at 25°C and 8.1×10^{-7} at 60°C . Calculate the Enthalpy Change for this reaction.

$$T_1 = 25^\circ\text{C} = 298\text{K} \quad K_1 = 3.6 \times 10^{-3}$$

$$T_2 = 333\text{K} \quad K_2 = 8.1 \times 10^{-7}$$

$$\Delta H^\circ = ?$$

$$\ln\left(\frac{K_2}{K_1}\right) = -\frac{\Delta H^\circ}{R} \left[\frac{1}{T_2} - \frac{1}{T_1} \right]$$

$$\Delta H^\circ = -\frac{R \ln\left(\frac{K_2}{K_1}\right)}{\frac{1}{T_2} - \frac{1}{T_1}} = \frac{-8.31 \ln\left(\frac{8.1 \times 10^{-7}}{3.6 \times 10^{-3}}\right)}{\frac{1}{333\text{K}} - \frac{1}{298\text{K}}}$$

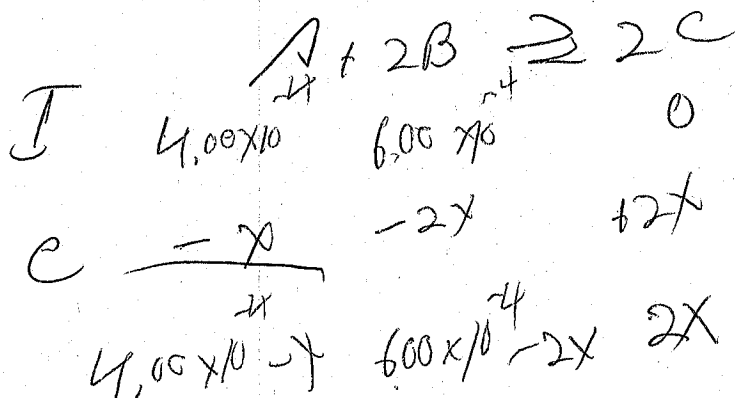
$$= -1.98 \times 10^3 \text{ J/mol} = -198 \text{ kJ/mol}$$

Chap 17 - Feb 14 - Beer-Lambert Supp HW: #S13

S13. Consider the aqueous solution equilibrium, $A(aq) + 2 B(aq) \rightleftharpoons 2 C(aq)$.
The product, C, has an absorption in the UV range of the spectrum at 320 nm, with a Molar Absorptivity, $\epsilon = 15,500 \text{ M}^{-1} \text{ cm}^{-1}$

A solution is prepared in a 0.50 cell with initial concentrations of A and B, $[A]_0 = 4.00 \times 10^{-4} \text{ M}$ and $[B]_0 = 6.00 \times 10^{-4} \text{ M}$, and the solution is allowed to reach equilibrium. At equilibrium, the percent transmission is $\%T = 32.0\%$.

Calculate the equilibrium constant, K_c , for this reaction.



[C] $A = \log \frac{100}{32} = \log \frac{100}{32} = 0.495$
 $\epsilon = 15,500 \text{ M}^{-1} \text{ cm}^{-1}$
 $b = 0.5 \text{ cm}$

[C] $A = \epsilon b c = \frac{0.495}{15,500 (0.5)}$
 $= 6.39 \times 10^{-5} \text{ M} = 2x$
 $x = 3.19 \times 10^{-5} \text{ M}$

[A] $= 4.00 \times 10^{-4} - 3.19 \times 10^{-5} = 3.68 \times 10^{-4}$
 [B] $= 6.00 \times 10^{-4} - 6.39 \times 10^{-5} = 5.36 \times 10^{-4}$

[C] $= 6.39 \times 10^{-5}$
 $K_c = \frac{[C]^2}{[A][B]^2} = 39$