Chap 17 - Feb 14 - Slide 59

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

To illustrate, consider the equilibrium: $2SO_3(g) \rightleftarrows 2SO_2(g) + O_2(g)$

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

$T_1 = 25^oC = 298K \quad K_1 = 3.6\times10^{-3}$

$T_2 = 333K \quad K_2 = 8.7\times10^{-7}$

$\Delta H^o = ?$

$$\frac{\ln(K_2/K_1)}{\frac{1}{T_2}-\frac{1}{T_1}} = -\frac{\Delta H^o}{R}$$

$$\Delta H = -\frac{R\ln(K_2/K_1)}{\frac{1}{T_2}-\frac{1}{T_1}} = \frac{-8.31\ln\frac{8.7\times10^{-7}}{3.6\times10^{-3}}}{\frac{1}{333K}-\frac{1}{298K}}$$

$= -1.98\times10^{3}\ \text{J/mol} = -198\ \text{kJ/mol}$

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, $\varepsilon = 15{,}500\ M^{-1}\ cm^{-1}$

A solution is prepared in a 0.50 cell with initial concentrations of A and B, $[A]_o = 4.00 \times 10^{-4}$ M and $[B]_o = 6.00 \times 10^{-4}$ 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.

$$A + 2B \rightleftharpoons 2C$$

	A	2B	2C
I	4.00×10^{-4}	6.00×10^{-4}	0
C	$-x$	$-2x$	$+2x$
	$4.00\times10^{-4} - x$	$6.00\times10^{-4} - 2x$	$2x$

[C]

$A = \log \frac{I_o}{I} = \log \frac{100}{32} = 0.495$

$\varepsilon = 15{,}500\ M^{-1}\ cm^{-1}$

$b = 0.5$ cm

$[C] = \frac{A}{\varepsilon b} = \frac{0.495}{15{,}500\,(0.5)}$

$= 6.39\times10^{-5}\ M = 2x$

$x = 3.19\times10^{-5}\ M$

$[A] = 4.00\times10^{-4} - 3.19\times10^{-5} = 3.68\times10^{-4}$

$[B]\ 6.00\times10^{-4} - 6.39\times10^{-5} = 5.36\times10^{-4}$

$[C] = 6.39\times10^{-5}$

$K_c = \frac{[C]^2}{[A][B]^2} = 39$