Marks 7

Calculate the partial pressure equilibrium constant, K_p, at 35 ºC for the reaction:

$$2NO_2(g) + Cl_2(g) \rightleftharpoons 2NOCl(g) + O_2(g)$$

In the reaction, 3 moles of gas react to form 3 moles of gas – there is no change in the number of moles of gas and so $\Delta n = 0$. K_c was calculated in the previous question (2005-N-5) to be 1.00×10^{-4}

$$K_p = K_c(RT)^{\Delta n}$$
$$= K_c(RT)^0 = K_c = 1.00\times10^4$$

K_p = **1.00×10^{-4}**

What is the standard free energy change, $\Delta G°$, for the forward reaction (in kJ mol^{-1}) at 35 ºC?

Using $\Delta G° = -RT\ln K_p$,

$$\Delta G° = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times ((35 + 273) \text{ K}) \times \ln(1.00 \times 10^{-4})$$
$$= -23600 \text{ J mol}^{-1} = -23.6 \text{ kJ mol}^{-1}$$

$\Delta G°$ = **-23.6 kJ mol^{-1}**

If 0.150 mol of $O_2(g)$ and 3.00×10^{-4} mol of $NO_2(g)$ are added to the 1.00 L flask, determine the free energy change, ΔG, (in kJ mol^{-1}) as the system moves to its new equilibrium point.

From the previous question (2005-N-5), the initial concentrations of [NOCl] and [Cl_2] are both equal to 0.0102 M. The concentrations of $O_2(g)$ and $NO_2(g)$ are equal to 0.150 M and 3.00×10^{-4} M as the flask has a volume of 1.00 L.

The reaction quotient, Q, for this reaction is $Q = \dfrac{[NOCl(g)]^2[O_2(g)]}{[NO_2(g)]^2[Cl_2(g)]}$. Hence:

$$Q = \frac{(0.0102)^2(0.150)}{(3.00\times10^{-4})^2(0.0102)} = 17000 \text{ and}$$

$$\Delta G = \Delta G^o + RT\ln Q = (-23.6\times10^3 \text{ J mol}^{-1}) + (8.314 \text{ J K}^{-1} \text{ mol}^{-1})\times((35+273)\text{ K})\times\ln(17000)$$
$$= +1.4 \text{ kJ mol}^{-1}$$

ΔG = **$+1.4$ kJ mol^{-1}**

Will the amount of $NO_2(g)$ in the flask increase or decrease as the system moves to its new equilibrium position? Explain.

As $Q > K_c$, the system will move to decrease Q. This occurs by reducing the concentration of the products and increasing the concentration of reactants. [$NO_2(g)$] will thus increase.