In Class Exercises Chapter 15

1.What is the equilibrium constant, K_c, given the data below. What is the K_p?

$$2BrCl_{(g)} \rightleftharpoons Br_{2(g)} + Cl_{2(g)}$$

[BrCl] = 0.500 atm, [Br_2] = 0.300 atm, [Cl_2] = 0.250 atm

$$K_p = \frac{\left(P_{Br_2}\right)\left(P_{Cl_2}\right)}{\left(P_{BrCl}\right)^2} = \frac{0.300\bullet 0.250}{0.500^2} = 0.300 \qquad 3\,pts$$

$$K_c = K_p/(RT)^{\Delta n} \quad 2\,pts$$

$$= 0.300/(0.0821\tfrac{atm\bullet L}{mol\bullet K}\times 298K)^{(2-2)} = 0.300 \quad 2\,pts$$

2. What is the equilibrium constant, K_c, given the data below.

$$2P_{(g)} + 3Cl_{(g)} \rightleftharpoons 2PCl_{3(g)} \qquad K_1 = 5$$

$$PCl_{3(g)} + Cl_{(g)} \rightleftharpoons PCl_{5(g)} \qquad K_2 = 0.25$$

$$2P_{(g)} + 5Cl_{(g)} \rightleftharpoons 2PCl_{5(g)} \qquad K_c = ?$$

$$2P_{(g)} + 3Cl_{(g)} \rightleftharpoons 2PCl_{3(g)} \qquad K_1 = 5$$

$$2\left(PCl_{3(g)} + Cl_{(g)} \rightleftharpoons PCl_{5(g)}\right) \quad K_2 = (0.25)^2$$

$$2P_{(g)} + 3Cl_{(g)} + \cancel{2PCl_{3(g)}} + 2Cl_{(g)} \rightleftharpoons \cancel{2PCl_{3(g)}} + PCl_{5(g)}$$

$$2P_{(g)} + 5Cl_{(g)} \rightleftharpoons 2PCl_{5(g)} \qquad K_c = K_1 \times K_2 = 5\times(0.25)^2 = 0.3$$

3. What are the equilibrium partial pressures of all the species in the system below if the initial NO_2 pressure is 0.250atm?

$$2NO_{2(g)} \rightleftharpoons 2NO_{(g)} + O_{2(g)} \qquad K_p = 6.5\times10^{-6}$$

	$NO_{2(g)}$	$NO_{(g)}$	$O_{2(g)}$
Initial	0.250	0.0	0.0
Change	-2x	+2x	+x
Eq	0.250 - 2x	+2x	+x

$$K_p = \frac{P^2_{NO_{(g)}} P_{O_{2(g)}}}{P^2_{NO_{2(g)}}} = \frac{4x^3}{(0.250-2x)^2} = 6.5\times10^{-6} \ \textit{(3pts)}$$

because we have a large concentration of $NO_{2(g)}$ and a small K_p we will try and assume 0.250 >> 2x

$$\frac{4x^3}{(0.250-2x)^2} \sim \frac{4x^3}{0.0625} = 6.5\times10^{-6} \rightarrow x = 4.67\times10^{-3}M$$

$$ck: \frac{2\left(4.67\times10^{-3}\right)}{0.250}\times100\% = 3.73\% < 5\%$$

P_{NO_2} = 0.250 - 2(4.67x10^{-3}) = 0.241 atm
P_{NO} = 2(4.67x10^{-3}) = 9.33x10^{-3}atm
P_{O_2} = 4.67x10^{-3} atm

4. Consider the following gas-phase reaction: $2A_{(g)} + B_{(g)} \rightleftharpoons C_{(g)} + D_{(g)}$. If the reaction starts out in equilibrium, in which direction will the reaction go to re-establish equilibrium if the system is subjected to the following changes:

(a) a decrease in volume

The side which has the largest number of gas moles will feel more concentrated. The reactants have 3 gas moles versus the products which have 2 gas moles so the reaction will proceed toward product to re-establish eq.

(b) an increase in temperature

This reaction is neither exothermic or endothermic so there is no effect.

(c) addition of reactants

Reaction shifts toward the right or forward or toward product

(d) addition of a catalyst

No effect so there is no shift from eq.

(e) addition of Ne

No effect so there is no shift from eq.

Furthermore, which of these changes affect the composition of the equilibrium but leave K_c unchanged? [(a) & (c)] Which changes affect the value of K_c? [(b)] Which affect neither the equilibrium composition nor K_c? [(d) & (e)]

5. At 700 K, K_p = 0.140 for the reaction below.

$$ClF_{3(g)} \rightleftharpoons ClF_{(g)} + F_{2(g)}$$

Calculate the equilibrium partial pressures of each species if the initial concentration of ClF_3 is 1.47 atm. What is the K_c at this temperature?

	$ClF_{3(g)}$	$ClF_{(g)}$	$F_{2(g)}$
Initial	1.47	0.0	0.0
Change	-x	+x	+x
Eq	1.47 - x	+x	+x

$$K_p = \frac{P_{ClF_{(g)}} P_{F_{2(g)}}}{P_{ClF_{3(g)}}} = \frac{x^2}{1.47 - x} = 0.140 \rightarrow x^2 - 0.140(1.47 - x) = 0$$

$$x^2 + 0.140x - 0.2058 = 0$$

$$for \quad ax^2 + bx + c = 0 \quad x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

in our case, a = 1, b = 0.140, c = -0.2058

$$x = \frac{-0.140 \pm \sqrt{(0.140)^2 - 4 \times 1 \times (-0.2058)}}{2 \times 1} = \frac{-0.140 \pm 0.918}{2}$$

$$x = 0.389\ atm \quad or \quad x = -0.529\ atm$$

$$P_{ClF_3} = 1.47 - 0.389 = 1.08\ atm$$

$$P_{ClF} = P_{F_2} = 0.389\ atm$$

$$ck: K_p = \frac{0.389^2}{1.08} = 0.140$$

For Kc:

$$K_p = K_c(RT)^{\Delta n} \rightarrow K_c = \frac{K_p}{(RT)^{\Delta n}} = \frac{0.140}{\left(0.0821\frac{L\bullet atm}{mol\cdot K}\times 700K\right)^1} = 2.44\times10^{-3}$$

6. For the reaction shown in problem 5, 9.25 g of ClF_3 was introduced into a 2.00 L container at 800K, 21.5% of ClF_3 decomposed to give an equilibrium mixture.

(a) What is the Kc?

$$[ClF_3] = 9.25\ g\ ClF_3 \times \frac{1\,mol}{92.453\ g\ ClF_3} \Big/ 2.00\ L = 0.0500M$$

$$[ClF_3]_{eq} = 0.0500M - \tfrac{21.5\%}{100\%}\times 0.0500M = \tfrac{100\%-21.5\%}{100\%}0.0500M = 0.0393M$$

$$[ClF]_{eq} = [F_2]_{eq} = \tfrac{21.5\%}{100\%}\times 0.0500M = 0.0108M$$

$$K_c = \frac{[ClF]_{eq}[F_2]_{eq}}{[ClF_3]_{eq}} = \frac{(0.0108M)^2}{0.0393M} = 2.94\times10^{-3}$$

(b) in a separate experiment, 39.4 g of ClF_3 was placed into a 2.00 L container at 800 K, what are the equilibrium concentrations of the species?

$$[ClF_3] = 39.4\ g\ ClF_3 \times \frac{1\,mol}{92.453\ g\ ClF_3} \Big/ 2.00\ L = 0.213M$$

	$ClF_{3(g)}$	$ClF_{(g)}$	$F_{2(g)}$
Initial	0.213	0.0	0.0
Change	-x	+x	+x
Eq	0.213 - x	+x	+x

$$K_c = \frac{[ClF][F_2]}{[ClF_3]} = \frac{x^2}{0.213-x} = 2.94\times10^{-3} \rightarrow x^2 + 0.00294x - 0.000626 = 0$$

$x = 0.0236M$ or $x = -0.0265M$

$$ck: K_c = \frac{0.0236^2}{0.213-0.0236} = 2.94\times10^{-3}$$

$$[ClF_3] = 0.213 - 0.0236 = 0.189M$$

$$[ClF] = [F_2] = 0.0236M$$