CHM 152/154 Diebolt Spring '05 **Quiz 2. Key**

1. A mixture of 1.79 g O_2 and 0.840 g N_2O is placed in a 5.00 L container at 50 °C. After equilibrium is established, there is 0.989 g of NO_2. (a) What are the equilibrium concentrations of O_2, N_2O and NO_2? (b) Calculate K_c. (c) Calculate K_p at 50 °C. (8 pts)

$$2N_2O(g) + 3O_2(g) \rightleftharpoons 4NO_2(g)$$

initial: $[N_2O] = \left(\frac{0.840\,g}{5.00\,L}\right)\left(\frac{1\,mol\,N_2O}{44.0\,g\,N_2O}\right) = 3.82\text{x}10^{-3}$ M; $[O_2] = \left(\frac{1.79\,g}{5.00\,L}\right)\left(\frac{1\,mol\,O_2}{32.0\,g\,O_2}\right) = 1.12\text{x}10^{-2}$ M

equilibrium: $[NO_2] = \left(\frac{0.989\,g}{5.00\,L}\right)\left(\frac{1\,mol\,NO_2}{46.0\,g\,NO_2}\right) = 4.30\text{x}10^{-3}$ M

	$[N_2O]$	$[O_2]$	$[NO_2]$
I, M	$3.82\text{x}10^{-3}$	$1.12\text{x}10^{-2}$	0
C	-2x	-3x	4x
E, M	$3.82\text{x}10^{-3}$ - 2x	$1.12\text{x}10^{-2}$ - 3x	4x=$4.30\text{x}10^{-3}$

$4x = 4.30\text{x}10^{-3} \Rightarrow x = 1.08\text{x}10^{-3}$

Eq concentrations: $[N_2O] = 3.82\text{x}10^{-3} - 2(1.08\text{x}10^{-3})$ = **$1.66\text{x}10^{-3}$ M N_2O**

$[O_2] = 1.12\text{x}10^{-2} - 3(1.08\text{x}10^{-3})$ = **$8.0\text{x}10^{-3}$ M O_2** (2 sig figs due to sig fig subtraction rule)

$[NO_2]$ = **$4.30\text{x}10^{-3}$ M NO_2**

b) $K_c = \frac{[NO_2]^4}{[N_2O]^2[O_2]^3} = \frac{[4.30x10^{-3}]^4}{[1.66x10^{-3}]^2[8.0x10^{-3}]^3}$ = **$2.4\text{x}10^2$**

c) $K_p = K_c(RT)^{\Delta n}$ $\Delta n = 4 - 5 = -1$ $K_p = 2.4\text{x}10^2(0.0821\times323)^{-1}$ = **9.1**

2. The gas arsine, AsH_3, decomposes by the following reaction:

$$2AsH_3(g) \rightleftharpoons 2As(s) + 3H_2(g)$$

In an experiment at a certain temperature, AsH_3 gas is placed in a flask at a pressure of 0.465 atm. After equilibrium has been established, the total pressure of the gases (for AsH_3 and H_2 taken together) is 0.579 atm. (a) What is the partial pressure of each gas at equilibrium? (b) Calculate the value of K_p for this reaction. (5 pts)

$AsH_3(g)$	$As(s)$	$H_2(g)$
0.465		0
-2x		+3x
0.465-2x		3x

Total P = P_{H_2} + P_{AsH_3}

0.579 = 0.465-2x + 3x

x = 0.114

a) P AsH_3 = 0.465-2(0.114) = **0.237 atm** P H_2 = 3x = 3(0.114) = **0.342 atm**

b) $K_p = \frac{P_{H_2}^{\ 3}}{P_{AsH_3}^{\ 2}} = \frac{(0.342)^3}{(0.237)^2}$ = **0.712**

3. Given the equations

$$6CH_4(g) \rightleftharpoons 3C_2H_6(g) + 3H_2(g) \qquad K_c = 8.6x10^{-37}$$

$$CH_4(g) + H_2O(g) \rightleftharpoons CH_3OH(g) + H_2(g) \qquad K_c = 2.8x10^{-21}$$

Calculate the value of K_c for: $2CH_3OH(g) + H_2(g) \rightleftharpoons C_2H_6(g) + 2H_2O(g)$
Make sure to show your work! (4 pt)

Divide the first equation by 3 ⇒ take K to the 1/3 power:

$$2CH_4(g) \rightleftharpoons C_2H_6(g) + H_2(g) \qquad K_c = (8.6x10^{-37})^{1/3} = 9.51x10^{-13}$$

Reverse & multiply the second equation by 2⇒ invert K and square K:

$$2CH_3OH(g) + 2H_2(g) \rightleftharpoons 2CH_4(g) + 2H_2O(g) \qquad K_c = \left[\frac{1}{2.8x10^{-21}}\right]^2 = 1.28x10^{41}$$

Add the two equations⇒multiply K values

$$2CH_3OH(g) + H_2(g) \rightleftharpoons C_2H_6(g) + 2H_2O(g) \qquad K_c = (9.51x10^{-13})(1.28x10^{41}) = \mathbf{1.2x10^{29}}$$

4. A mixture of 0.500 atm H_2 and 0.500 atm CO_2 is placed in a container and undergoes the following reaction:

$$H_2(g) + CO_2(g) \rightleftharpoons CO(g) + H_2O(g) \qquad K_p = 0.771$$

Calculate the equilibrium partial pressures of each of the four gases. (5 pts)

	H_2	CO_2	CO	H_2O
I (atm)	0.500	0.500	0	0
C	-x	-x	+x	+x
E (atm)	0.500 - x	0.500 - x	x	x

$$K_p = \frac{P_{H_2O}P_{CO}}{P_{H_2}P_{CO_2}} \qquad 0.771 = \frac{x \cdot x}{(0.500 - x)(0.500 - x)} = \frac{x^2}{(0.500 - x)^2}$$

$$0.878 = \frac{x}{0.500 - x} \quad \Rightarrow \quad 0.878(0.500 - x) = x$$

0.439-0.878x = x

0.439 = 1.878 x

x = 0.439/1.878 ⇒ **x = 0.234**

Equilibrium concentrations: $\mathbf{P_{CO} = P_{H_2O} = 0.234\ atm}$

$\mathbf{P_{CO_2} = P_{H_2} = 0.500 - 0.234 = 0.266\ atm}$

5. The value of K_c for the following reaction is 3.17 at 300 K.

$XeF_2(g) + F_2(g) \rightleftharpoons XeF_4(g)$

Suppose 0.525 moles of XeF_2 and 1.12 moles of F_2 are placed in a 2.50 L vessel. What are the equilibrium concentrations of XeF_2, F_2, and XeF_4? (8 pts)

Initial: $[XeF_2] = \frac{0.525\,moles}{2.50\,L}$ = 0.210 M $\quad$ $[F_2] = \frac{1.12\ moles}{2.50\,L}$ = 0.448 M

	$[XeF_2]$	$[F_2]$	$[XeF_4]$
I,	0.210	0.448	0
C	-x	-x	+x
E	0.210 - x	0.448 - x	x

$K_c = \frac{[XeF_4]}{[XeF_2][F_2]}$ $\quad$ $3.17 = \frac{x}{(0.210 - x)(0.448 - x)}$

3.17(0.210 – x)(0.448 – x) = x

3.17[0.09408 – 0.658x + x^2] = x

0.298 – 2.086x + $3.17x^2$ = x

$3.17x^2$ – 3.086x + 0.298 = 0

$$x = \frac{+3.086 \pm \sqrt{(3.086)^2 - 4(3.17)(0.298)}}{2(3.17)} = \frac{+3.086 \pm 2.397}{6.34}$$

x = 0.865 or 0.109 so **x = 0.109 M; x can't be 0.865 or we get (-) eq concentrations**

Equilibrium concentrations:

$[XeF_2]$ = 0.210 – 0.109 = 0.101 M XeF_2

$[F_2]$ = 0.448 – 0.109 = 0.339 M F_2

$[XeF_4]$ = 0.109 M XeF_4