NAME:  

CHM 2046 Quiz 1B Spring, 2017

Answer all questions. Give your final answer with correct units, if any, and to the correct sig. figs. Useful Information: 0 ºC ≈ 273 K, R = 0.0820 L. atm/mol.K

1) a) (3 points each) For the following equilibria, write down their mass-action expression, $Q_c$.

- $2H_2O_2 (g) + 2NO (g) \rightleftharpoons 2H_2O (g) + N_2O_4 (g)$
- $P_4(s) + 6Cl_2(g) \rightleftharpoons 4PCl_3(l)$

$$Q_c = \frac{[N_2O_4][H_2O]^2}{[NO]^2[H_2O_2]^2}$$

$$Q_c = \frac{1}{[Cl_2]^6}$$

b) (5 points) At a particular temperature, $K_p = 313$ for the reaction

$$2 SO_2 (g) + O_2 (g) \rightleftharpoons 2 SO_3 (g)$$

In a particular reaction carried out, the pressures of the gases at a particular time are 0.12 atm (atmospheres) of $SO_2$, 1.2 atm of $O_2$, and 2.4 atm of $SO_3$. Has the reaction reached equilibrium yet? Explain.

If not, in which direction is it proceeding? Explain.

$$Q_p = \frac{p(SO_3)^2}{p(SO_2)^2p(O_2)} = \frac{(2.4)^2}{(0.12)^2(1.2)^2} = 333 \quad \therefore Q_p < K_p \quad \therefore \text{reaction proceeding to the left.}$$

c) (4 points) $N_2O_4 (g)$ was introduced into a 2.00 L flask and allowed to reach equilibrium at 100 ºC. At equilibrium, the flask contained 0.38 mol of $N_2O_4$ and 0.40 mol of $NO_2$. What is $K_c$ for this reaction at 100 ºC?

$$K_c = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(0.20)^2}{0.19} = 0.21$$

2) (6 points total)

a) At 350 K, the reaction below has $K_c = 1.05$. What is $K_p$? Explain.

$$2 CH_2Cl_2 (g) \rightleftharpoons CH_4 (g) + CCl_4 (g)$$

$$K_p = K_c$$

$$K_p = 1.05$$

b) At 300 K, the reaction below has $K_c = 279$. What is $K_p$? Explain.

$$2 SO_2 (g) + O_2 (g) \rightleftharpoons 2 SO_3 (g)$$

$$\Delta n = -1 \quad \therefore$$

$$K_p = \frac{K_c}{RT} = \frac{279}{(0.0820)(300)} = 11.3$$

$$K_p = 11.3$$

c) At 300 K, what is $K_c$ for the reaction below? Explain.

$$4 SO_3 (g) = 4 SO_2 (g) + 2 O_2 (g)$$

This is the reaction of part b) reversed and multiplied by 2.

$$\therefore \text{new } K_c = \frac{1}{K_c^2} = \frac{1}{279^2} = 0.0000128 = 1.28 \times 10^{-5}$$
4. a) (6 points) \[ 2\text{HBr (g)} \rightleftharpoons \text{H}_2 (g) + \text{Br}_2 (g) \]

A 2.00 L flask is filled with 0.300 mol of HBr and allowed to reach equilibrium at a particular temperature. At equilibrium, \([\text{HBr}] = 0.104 \text{ M}\). Calculate \(K_c\) at this temperature.

\[ 2 \text{HBr} \rightleftharpoons \text{H}_2 + \text{Br}_2 \]
\[
\begin{array}{c|ccc}
\text{I} & 0.150 & 0 & 0 \\
\text{C} & -2x & +x & +x \\
\text{E} & (0.150-2x) & x & x \\
\hline
\end{array}
\]

\[(0.150-2x) = 0.104\]
\[
\therefore x = 0.023
\]

\[ K_c = \frac{[\text{H}_2][\text{Br}_2]}{[\text{HBr}]^2} = \frac{x^2}{(0.15 - 2x)^2} = \frac{(0.023)^2}{(0.104)^2} \]
\[
\therefore K_c = 4.89 \times 10^{-2}
\]

b) (8 points) \[ \text{CO}_2 (g) + \text{H}_2 (g) \rightleftharpoons \text{CO} (g) + \text{H}_2O (g) \quad (K_c = 2.4) \]

0.500 mol each of \(\text{CO}_2\) and \(\text{H}_2\) are placed in a 10.0 L flask and allowed to reach equilibrium at some temperature. Calculate the equilibrium concentrations of all species.

\[ \text{CO}_2 + \text{H}_2 \rightleftharpoons \text{CO} + \text{H}_2O \]
\[
\begin{array}{c|ccc}
\text{I} & 0.0500 & 0.0500 & 0 & 0 \\
\text{C} & -x & -x & +x & +x \\
\text{E} & (0.0500-x) (0.0500-x) & x & x \\
\hline
\end{array}
\]

\[ K_c = 2.4 = \frac{x^2}{(0.0500-x)^2} \]
\[
\therefore \pm 1.549 = \frac{x}{0.0500-x} \quad \therefore x = 0.030
\]

\[ 1.549x + 0.077 = x \quad [\text{CO}_2] = [\text{H}_2] = 0.020 \]
\[ 2.549x = 0.077 \quad [\text{CO}] = [\text{H}_2O] = 0.030
\]

5. (14 points) Consider the reaction in Question 4b again.

\[ \text{CO}_2 (g) + \text{H}_2 (g) \rightleftharpoons \text{CO} (g) + \text{H}_2O (g) \quad (K_c = 2.4) \]

This time 0.500 mol of \(\text{CO}_2\) and 1.00 mol of \(\text{H}_2\) are placed in a 10.0 L flask and allowed to reach equilibrium at some temperature. Calculate the equilibrium concentrations of all species.

\[ \text{CO}_2 + \text{H}_2 \rightleftharpoons \text{CO} + \text{H}_2O \]
\[
\begin{array}{c|ccc}
\text{I} & 0.0500 & 0.100 & 0 & 0 \\
\text{C} & -x & -x & +x & +x \\
\text{E} & (0.0500-x) (0.100-x) & x & x \\
\hline
\end{array}
\]

\[ K_c = 2.4 = \frac{x^2}{(0.0500-x)(0.100-x)} \]
\[
\therefore 2.4x^2 - 0.36x + 0.012 = x^2
\]
\[ 1.4x^2 - 0.36x + 0.012 = 0 \]
\[ x = \frac{0.36 \pm \sqrt{0.1296 - 4(1.4)(0.012)}}{2.8} = \frac{0.36 \pm \sqrt{0.0624}}{2.8} \]
\[ x = \frac{0.36 \pm 0.250}{2.8} = 0.039 \text{ or } 0.218
\]

We choose \(x = 0.039\)

\[ [\text{CO}_2] = 0.011\text{M} \]
\[ [\text{H}_2] = 0.061\text{M} \]
\[ [\text{CO}] = [\text{H}_2O] = 0.039\text{M} \]