CHEM 130 **Name ___________________________________**
Quiz 9 – November 30, 2011 **Due In Class Monday December 5, 2011.**

5 Point Bonus for all Quizzes Submitted by 4:00 PM Friday, December 2

Complete the following individually. You may use your textbook and notes, but may not receive assistance from your classmates or anyone other than Dr. Lamp. *This signed statement must accompany the completed assignment*. By signing below, you certify that you completed the problems in accordance with these rules. No credit will be given to unsigned papers. Staple any additional sheets prior to turning the assignment in.

Signature____ Date________________

Complete the following problems on separate paper and **staple the pages** to this sheet **write your final answers on this page**. **You must show your work to receive full credit**. Show your answers to the correct number of significant figures with the correct units.

1. A 0.431 g sample of HCl(g) was placed in a 625 mL reaction vessel at a 862K and allowed to dissociate to H_2 and Cl_2. When equilibrium is reached between HCl(g), H_2(g) and Cl_2(g), 0.0414 g Cl_2 is present. What is the K_c for the reaction at this temperature?

Answer: K_c = 0.00300_3

First we need a balanced reaction: $2HCl \rightleftarrows H_2 + Cl_2$. To determine the equilibrium constant, we need values for equilibrium concentrations for each of the species. We are given information related to the initial amount of HCl and the equilibrium amount of Cl_2. One approach is to treat this with the ICE table. To do so, we need to convert the masses into molarities.

$$0.431 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.461 \text{ g}} \times \frac{1}{0.625 \text{ L}} = 0.0189_1 \text{ M HCl}$$

$$0.0414 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.9054 \text{ g}} \times \frac{1}{0.625 \text{ L}} = 9.34_2 \times 10^{-4} \text{ M } Cl_2$$

	$2HCl$	$\rightleftarrows$	H_2	+	Cl_2
I	0.0189_1		0		0
C	$-2(9.34_2 \times 10^{-4})$		$9.34_2 \times 10^{-4}$		$9.34_2 \times 10^{-4}$
E	$0.0189_1 - 2(9.34_2 \times 10^{-4})$ $= 0.0170_4$M		$9.34_2 \times 10^{-4}$ M		$9.34_2 \times 10^{-4}$ M

Therefore, K_c is:

$$K_c = \frac{[H_2][Cl_2]}{[HCl]^2} = \frac{(9.34_2 \times 10^{-4} \text{ M})(9.34_2 \times 10^{-4} \text{ M})}{(0.0170_4 \text{M})^2} = 0.00300_3$$

2. Carbonyl bromide, $COBr_2(g)$, decomposes to $CO(g)$ and $Br_2(g)$ with an equilibrium constant, K_c, of 0.190 at 73°C. If a 0.0150 mol sample of $COBr_2$ is heated in a 2.50 L flask until equilibrium is attained, what will be the concentrations of all $COBr_2$, CO, and Br_2 at equilibrium?

Answer: $[COBr_2]$ = 0.000178 M

[CO] = 0.00582 M

$[Br_2]$ = 0.00582 M

The ICE table approach works well here. First we need to get things in terms of concentration:

$$\frac{0.0150 \text{ mol } Br_2}{2.5\text{L}} = 0.00600 \text{ M } COBr_2$$

	$COBr_2$	$\rightleftarrows$	CO	+	Br_2
I	0.00600		0		0
C	-x		+x		+x
E	0.00600-x		x		x

Inserting into K_c gives:

$$K_c = \frac{[CO][Br_2]}{[COBr_2]} = \frac{(x)(x)}{0.00600-x}$$

Now some algebra:

$$(0.00600-x)K_c = x^2$$
$$0 = x^2 + K_c x - 0.00600K_c = (0.190)^2 + 0.190x - 0.00114$$

From the quadratic formula, we find $x = 0.00582_2$ or -0.196

Since x represents the equilibrium concentration of CO and Br_2, a negative value makes no chemical sense, therefore, the value $x = 0.00582_2$ is the reasonable result.

Now to find the equilibrium concentrations of all species:

$[CO] = [Br_2] = x = 0.00582$ M

$[COBr_2] = 0.00600 - x = 0.000178$ M

3. In the reaction $CO(g) + H_2O(g) \rightleftarrows CO_2(g) + H_2(g)$, $K_c = 31.4$ at 588 K. If 10.1 g of each reactant and product are brought together in a 1.00 L reaction vessel at 588 K, how many grams of H_2O will be present at equilibrium?

Answer: mass H_2O = 6.83_1 g H_2O

The ICE table approach works well here. First we need to get things in terms of concentration:

$$10.1 \text{ g CO} \times \frac{1 \text{ mol CO}}{28.010 \text{ g}} \times \frac{1}{1 \text{ L}} = 0.360_6 \text{ M CO}$$

Repeating the process for the other reactants and products we find the following initial concentrations:
$[H_2O] = 0.560_6$ M, $[CO_2] = 0.229_4$ M, $[H_2] = 5.01_0$ M

Now we need to figure out what direction the reaction will go on the way to equilibrium:

$$Q = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{(0.229_4)(5.01_0)}{(0.360_6)(0.560_6)} = 5.69$$

Since Q is less than K_c, we need to make more products and the reaction will proceed to the right.

	CO	+	H_2O	$\rightleftarrows$	CO_2	+	H_2
I	0.360_6		0.560_6		0.229_4		5.01_0
C	-x		-x		+x		+x
E	0.360_6-x		0.560_6-x		0.229_4+x		5.01_0+x

Inserting into K_c gives:

$$K_c = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{(0.229_4+x)(5.01_0+x)}{(0.360_6-x)(0.560_6-x)} = \frac{1.14_9 + 5.23_9x + x^2}{0.202_1 - 0.921_2x + x^2}$$

Now some algebra:

$$(0.202_1 - 0.921_2x + x^2)K_c = 1.14_9 + 5.23_9x + x^2$$
$$6.34_8 - 28.9_2x + 31.4x^2 = 1.14_9 + 5.23_9x + x^2$$
$$30.4x^2 - 34.1_6x + 5.19_8 = 0$$

From the quadratic formula, we find $x = 0.942_2$ or 0.181_4
If we use the value $x = 0.942_2$ M, we will arrive at negative concentrations for CO and H_2O, which makes no chemical sense. Therefore, $x = 0.181_4$ M is the reasonable result.

Thus, the equilibrium concentration of H_2O is $0.560_6 - 0.181_4 = 0.379_2$ M
And the mass of H_2O is:

$$\frac{0.379_2 \text{ mol } H_2O}{1 \text{ L}} \times 1.00 \text{ L} \times \frac{18.015 \text{ g } H_2O}{1 \text{ mol } H_2O} = 6.83_1 \text{ g } H_2O$$

4. Determine K_c for the reaction: $N_2(g) + O_2(g) + Cl_2(g) \rightleftharpoons 2NOCl(g)$ from the following data at 298K:

$\frac{1}{2} N_2(g) + O_2(g) \rightleftharpoons NO_2(g)$	$K_p = 1.0 \times 10^{-9}$
$NOCl(g) + \frac{1}{2} O_2(g) \rightleftharpoons NO_2Cl(g)$	$K_p = 1.1 \times 10^{2}$
$NO_2(g) + \frac{1}{2} Cl_2(g) \rightleftharpoons NO_2Cl(g)$	$K_p = 0.3$

Answer: K_c = 1.8×10^{-22}

We need to rearrange reactions to make our target reaction:

$2(\frac{1}{2} N_2(g) + O_2(g) \rightleftharpoons NO_2(g))$	$K_p = (1.0 \times 10^{-9})^2$
$2(NO_2Cl(g) \rightleftharpoons NOCl(g) + \frac{1}{2} O_2(g))$	$K_p = 1/(1.1 \times 10^{2})^2$
$2(NO_2(g) + \frac{1}{2} Cl_2(g) \rightleftharpoons NO_2Cl(g))$	$K_p = (0.3)^2$

So, the sum of the reactions is:

$$N_2 + \cancel{2}O_2 + 2\cancel{NO_2Cl} + \cancel{2NO_2} + Cl_2 \rightleftharpoons \cancel{2NO_2} + \cancel{2NOCl} + \cancel{O_2} + 2NO_2Cl$$

$$N_2 + O_2 + Cl_2 \rightleftharpoons 2NO_2Cl$$

$$K_p = \frac{(1.0 \times 10^{-9})^2(0.3)^2}{(1.1 \times 10^{2})^2} = 7.4 \times 10^{-24}$$

Now we need to convert from K_p to K_c:

$$K_c = K_p/(RT)^{\Delta n}$$

In this reaction we have 2 moles of gaseous products and 1 mole of gaseous reactions, so $\Delta n = -1$

$K_c = (7.4 \times 10^{-24})/(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}\ 298\text{K})^{-1}$ = **1.8×10^{-22}**