Mr. Bracken
AP Chemistry

Name BRACKEN KEY
Period ____

Equilibrium Worksheet #5

From the Internet October, 2003 (Revised 2010)

1. Suppose 1.000 mole CO and 3.000 moles H_2 are put into a 10.00 L vessel at 1200 Kelvin. If the equilibrium constant for the reaction shown below is 3.92, calculate the final composition of the mixture at equilibrium.

	$CO_{(g)}$	+	$3H_{2(g)}$	$\leftrightarrows$	$CH_{4(g)}$	+	$H_2O_{(g)}$
R							
I	.1000 M		.3000 M		0		0
C	$-x$		$-3x$		$+x$		$+x$
E	.1000 $-x$		.3000 $-3x$		x		x

[CO] = 0.0613 M
[H_2] = 0.1839 M
[CH_4] = 0.0387 M
[H_2O] = 0.0387 M

$$3.92 = \frac{[x][x]}{[.1000-x][.3000-3x]^3}$$

$\boxed{\text{Guess} = .01 \quad x = .0387}$

$[CO] = .1000 - .0387 = .0613\ M$
$[H_2] = .3000 - (3)(.0387) = .1839\ M$
$[CH_4] = .0387\ M$
$[H_2O] = .0387\ M$

2. For the reaction represented by the equation:

$$H_{2(g)} + CO_{2(g)} \leftrightarrows H_2O_{(g)} + CO_{(g)}$$

the value of K is 0.771 at 750°C. If 0.0100 mole of $H_{2(g)}$ and 0.0100 mole of $CO_{2(g)}$ are mixed inside a 2.00 L vessel at 750°C, what are the concentrations of all substances at equilibrium?

	$H_{2(g)}$	+	$CO_{2(g)}$	$\leftrightarrows$	$H_2O_{(g)}$	+	$CO_{(g)}$
R							
I	.00500 M		.00500 M		0		0
C	$-x$		$-x$		$+x$		$+x$
E	.00500 $-x$		.00500 $-x$		x		x

K = .771

[H_2] = 0.00266 M
[CO_2] = 0.00266 M
[H_2O] = 0.00234 M
[CO] = 0.00234 M

$$.771 = \frac{[x][x]}{[.00500-x][.00500-x]}$$

$\boxed{x = .00234}$

$[H_2O] = .00234\ M$
$[CO] = .00234\ M$
$[H_2] = .00500 - .00234 = .00266$
$[CO_2] = .00500 - .00234 = .00266\ M$

3. Nitrogen monoxide is formed in automobile exhaust by the reaction of the N_2 and O_2 in the air. At 2127°C, K=0.0125. Initially, a mixture contains 0.850 moles of N_2 and 0.850 moles of O_2 in a 15.0-Liter vessel. Calculate the concentration of all species when equilibrium is established at 2127°C.

	N_2 (g)	+	O_2 (g)	$\leftrightarrows$	2 NO (g)
R					
I	.0567 M		.0567 M		0
C	$-x$		$-x$		$+2x$
E	.0567 $-x$		.0567 $-x$		$2x$

[N_2] = 0.0537 M
[O_2] = 0.0537 M
[NO] = 0.00600 M

$$K = .0125 = \frac{[2x]^2}{[.0567-x][.0567-x]}$$

$\boxed{x = .00301}$

Then plug into "E" line to obtain final molarities

4. The equilibrium constant K is 115 at 60°C for the reaction shown below. A 1.500 L flask contains 3.000 mole of each substance.

	$F_{2\,(g)}$	+	$H_{2\,(g)}$	$\leftrightarrows$	$2HF_{(g)}$	$K = 115$
R						
I	2.000M		2.000M		2.000M	
C	$-x$		$-x$		$+2x$	
E	$2.000-x$		$2.000-x$		$2.000+2x$	

A. Is the system at equilibrium? (Show work here)

$$Q = \frac{[2.000]^2}{[2.000][2.000]} = 1$$

$K > Q$

NOT AT EQUILIBRIUM SINCE $K \neq Q$

B. If not, in which direction will the equilibrium shift?

RIGHT

C. What are the equilibrium concentrations for each substance?

$[H_2] = 0.472$ M
$[F_2] = 0.472$ M
$[HF] = 5.056$ M

$$115 = \frac{[2.000 + 2x]^2}{[2.000 - x][2.000 - x]}$$

$$\boxed{x = 1.53}$$

$[F_2] = 2.000 - 1.53 = .47$ M

$[H_2] = 2.000 - 1.53 = .47$ M

$[HF] = 2.000 + 2(1.53) = 5.06$ M

5. The value of K for the HI equilibrium at 425°C is 54.8.

$$H_{2\,(g)} + I_{2\,(g)} \leftrightarrows 2HI_{(g)} \qquad K = 54.8$$

A quantity of $HI_{(g)}$ is placed in a 1.00 L container and allowed to come to equilibrium at 425°C. At equilibrium, the concentration of $HI_{(g)}$ is found to be 0.50 M.

(a) What are the concentrations of $H_{2\,(g)}$ and $I_{2\,(g)}$ at equilibrium?
(b) What was the initial concentration of $HI_{(g)}$?

$y - 2x = .50$

$H_{2\,(g)}$	+	$I_{2\,(g)}$	$\leftrightarrows$	$2HI_{(g)}$
0		0		y
$+x$		$+x$		$-2x$
x		x		$y-2x$

$[H_2] = 0.068$ M
$[I_2] = 0.068$ M
$[HI] = 0.64$ M

$$K = 54.8 = \frac{[y-2x]^2}{[x][x]}$$

$$54.8 = \frac{[.50]^2}{[x][x]}$$

$$\boxed{x = .068}$$

(a) $[H_2] = [I_2] = .068$ M

$[HI] = y - 2(.068) = .50$

$y - .136 = .50$

$\boxed{y = .636\text{ M}}$

6. Consider the system $A_{(g)} + B_{(g)} \leftrightarrows AB_{(g)}$ at equilibrium where $K_c = 500.$

A B AB

At equilibrium, the concentrations of A, B, and AB are found to be 0.0300 M, 0.0100 M, and 0.150 M, respectively, in a 5.00 L container. An additional 0.0400 moles of B are added. What are the final equilibrium concentrations of A, B, and AB?

$\frac{.0400}{5.00} = .008M$ added to B for initial molarity of B

[A] = 0.0246 M
[B] = 0.0126 M
[AB] = 0.1554 M

$A_{(g)}$	+ $B_{(g)}$	$\leftrightarrows AB_{(g)}$
.0300 M	.0100 + .008 = .0180M	.150 M
$-x$	$-x$	$+x$
$.0300-x$	$.0180-x$	$.150+x$

$$K = 500 = \frac{[.150+x]}{[.0300-x][.0180-x]}$$

Solver Guess = .01
x = .0054

$[A] = .0300 - .0054 = .0246M$

$[B] = .0180 - .0054 = .0126M$

$[AB] = .150 + .0054 = .1554M$

7. A 2.50-mole quantity of NOCl was initially in a 1.50-L reaction chamber at 400°C where the following chemical reaction occurred.

$$2NOCl\ (g) \leftrightarrows 2\ NO\ (g) + Cl_2\ (g)$$

After equilibrium was established, it was found that 28.0% of the NOCl had decomposed. Calculate the equilibrium constant K_c for this reaction.

K = 0.0353

molarity $= \frac{2.50}{1.50} = 1.67\ M$

28% decomposes or breaks down

$1.67 \times .28 = .467\ M$ changes from original 1.67M NOCl

$2x = .467$

$x = .233$

$2\ NOCl$	$\rightleftarrows 2\ NO$	$+ Cl_2$
1.67 M	0	0
$-2x$	$+2x$	$+x$
$1.67-2x$	$2x$	x

$$K = \frac{[2x]^2[x]}{[1.67-2x]^2} \qquad K = \frac{[.467]^2[.233]}{[1.67-.467]^2} = .0351$$