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AP Chemistry

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Equilibrium Worksheet #4
From the Internet October, 2003 (Revised 2011)

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.

$CO_{(g)}$	+	$3H_{2(g)}$	⇆	$CH_{4(g)}$	+	$H_2O_{(g)}$
.1000 M		.3000 M		0		0
− x		−3x		+ x		+ x
.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

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

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?

$H_{2(g)}$	+	$CO_{2(g)}$	⇆	$H_2O_{(g)}$	+	$CO_{(g)}$
.005 M		.005 M		0		0
− x		− x		+ x		+ x
.005 − x		.005 − x		x		x

K = .771

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

$$K = .771 = \frac{[x][x]}{[.005-x][.005-x]}$$

x = .00234

[H_2] = .005 − .00234 = .00266 M
[CO_2] = .005 − .00234 = .00266 M
[H_2O] = .00234 M
[CO] = .00234 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)	⇆	2 NO (g)
.0567 M		.0567 M		0
− x		− x		+ 2x
.0567 − x		.0567 − x		2x

K = .0125

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

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

x = .00300

[N_2] = .0567 − .00300 = .0537 M
[O_2] = .0567 − .00300 = .0537 M
[NO] = (2)(.00300) = .00600 M

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.

$$F_{2(g)} + H_{2(g)} \leftrightharpoons 2HF_{(g)} \quad K = 115$$

$F_{2(g)}$	$H_{2(g)}$	$2HF_{(g)}$
2.000 M	2.000 M	2.000 M
−x	−x	+ 2x
2−x	2−x	2+2x

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

$Q = \frac{[2]^2}{[2][2]} = 1$ $K > Q$ System NOT @ equilibrium

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

C. What are the equilibrium concentrations for each substance?

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

$x = 1.528$

$[F_2] = 2 - 1.528 = .472\ M$

$[H_2] = 2 - 1.528 = .472\ M$

$[HF] = 2 + (2)(1.528) = 5.056\ M$

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

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

$$H_{2(g)} + I_{2(g)} \leftrightharpoons 2HI_{(g)}$$

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 $HI_{(g)}$?

$H_{2(g)}$	$I_{2(g)}$	$2HI_{(g)}$
0	0	y
+x	+x	−2x
x	x	y−2x = .50

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

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

$x = .068$

AT EQUILIBRIUM
$[H_2] = .068\ M$
$[I_2] = .068\ M$

$y - 2(.068) = .50$
$y - .136 = .50$
$y = .636$
$y = .64\ M$

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

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?

$\frac{.0400}{5.00} = .008$ M B added

$A_{(g)}$	+ $B_{(g)}$	$\leftrightarrows AB_{(g)}$
.0300M	.018 M	.150 M
−x	−x	+x
.0300−x	.018−x	.150+x

$Q = \frac{[.150]}{[.0300][.018]}$

Q = 278

K > Q

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

$K = 500 = \frac{[.150+x]}{[.0300-x][.018-x]}$

x = .00538

[A] = .0300 − .00538 = .0246 M
[B] = .018 − .00538 = .0126 M
[AB] = .150 + .00538 = .1554 M

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

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

2 NOCl	$\rightleftharpoons$ 2 NO	+ Cl_2
1.67 M	0	0
−2x	+2x	+x
1.67 −2x	2x	x

(1.67)(.28) = .468 M has decomposed

Since 28.0% of NOCl has decomposed, then we calculate the molarity change for NOCl is .468 M. This number represents 2x in the RICE chart. 2x = .468

x = .234

$$K = \frac{[2x]^2[x]}{[1.67-2x]^2} = \frac{[(2)(.234)]^2[.234]}{[1.67-2(.234)]^2} = \frac{.0513}{1.44} = .0356$$