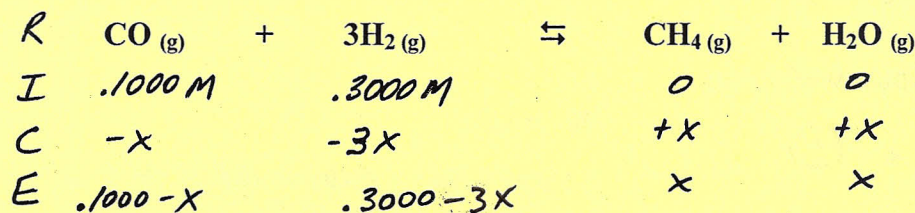


Equilibrium Worksheet #5
From the Internet October, 2003 (Revised 2010)

1. Suppose 1.000 mole CO and 3.000 moles H₂ 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] = 0.0613 M
[H₂] = 0.1839 M
[CH₄] = 0.0387 M
[H₂O] = 0.0387 M

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

Guess = .01

x = .0387

[CO] = .1000 - .0387 = .0613 M

[H₂] = .3000 - (3)(.0387) = .1839 M

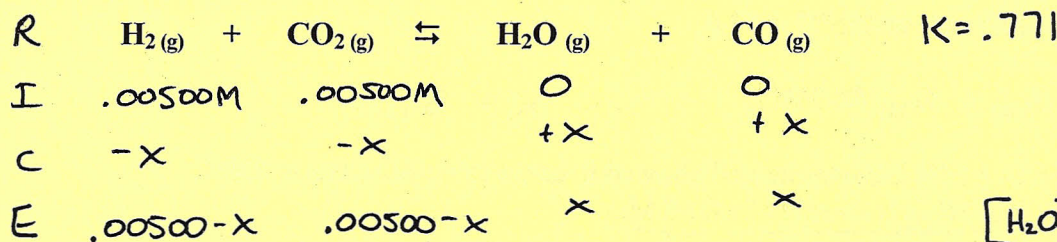
[CH₄] = .0387 M

[H₂O] = .0387 M

2. For the reaction represented by the equation:



the value of K is 0.771 at 750°C. If 0.0100 mole of H₂(g) and 0.0100 mole of CO₂(g) are mixed inside a 2.00 L vessel at 750°C, what are the concentrations of all substances at equilibrium?



[H₂] = 0.00266 M
[CO₂] = 0.00266 M
[H₂O] = 0.00234 M
[CO] = 0.00234 M

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

x = .00234

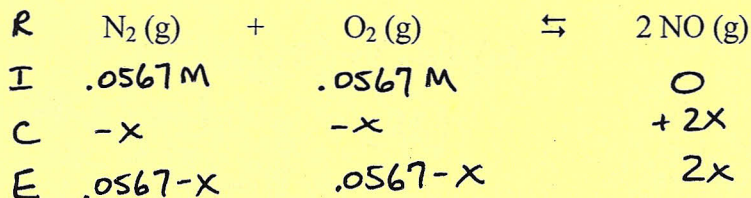
[H₂O] = .00234 M

[CO] = .00234 M

[H₂] = .00500 - .00234 = .00266 M

[CO₂] = .00500 - .00234 = .00266 M

3. Nitrogen monoxide is formed in automobile exhaust by the reaction of the N₂ and O₂ in the air. At 2127°C, K=0.0125. Initially, a mixture contains 0.850 moles of N₂ and 0.850 moles of O₂ in a 15.0-Liter vessel. Calculate the concentration of all species when equilibrium is established at 2127°C.



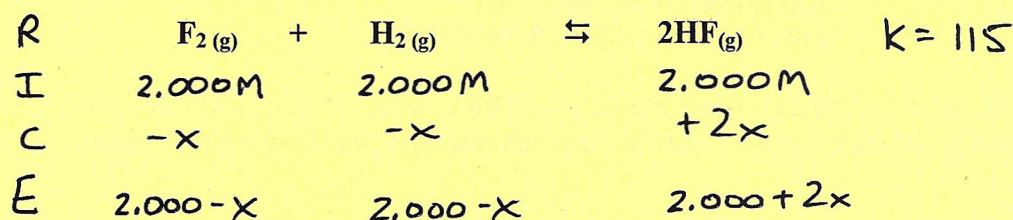
[N₂] = 0.0537 M
[O₂] = 0.0537 M
[NO] = 0.00600 M

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

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.



- 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?

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

$$\boxed{x = 1.53}$$

$$[\text{F}_2] = 2.000 - 1.53 = .47 \text{ M}$$

$$[\text{H}_2] = 2.000 - 1.53 = .47 \text{ M}$$

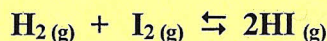
$$[\text{HF}] = 2.000 + 2(1.53) = 5.06 \text{ M}$$

$$[\text{H}_2] = 0.472 \text{ M}$$

$$[\text{F}_2] = 0.472 \text{ M}$$

$$[\text{HF}] = 5.056 \text{ M}$$

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

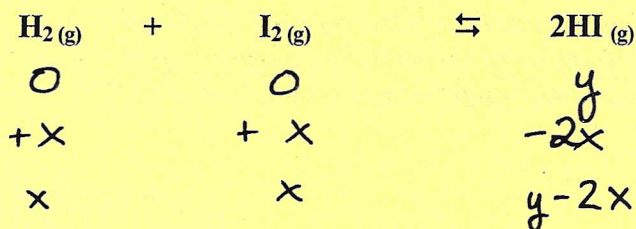


$$K = 54.8$$

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

- (a) What are the concentrations of $\text{H}_2(\text{g})$ and $\text{I}_2(\text{g})$ at equilibrium?
(b) What was the initial concentration of $\text{HI}(\text{g})$?

$$y - 2x = .50$$



$$[\text{H}_2] = 0.068 \text{ M}$$

$$[\text{I}_2] = 0.068 \text{ M}$$

$$[\text{HI}] = 0.64 \text{ M}$$

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

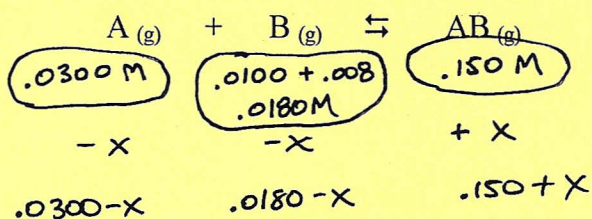
$$54.8 = \frac{[.50]^2}{[x][x]}$$

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

$$\begin{aligned} \text{(a)} \quad [\text{H}_2] &= [\text{I}_2] = .068 \text{ M} \\ [\text{HI}] &= y - 2(.068) = .50 \\ y - .136 &= .50 \\ \boxed{y} &= .636 \text{ M} \end{aligned}$$

6. Consider the system $A_{(g)} + B_{(g)} \rightleftharpoons AB_{(g)}$ 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?



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

$$\begin{aligned}
 [A] &= 0.0246 \text{ M} \\
 [B] &= 0.0126 \text{ M} \\
 [AB] &= 0.1554 \text{ M}
 \end{aligned}$$

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

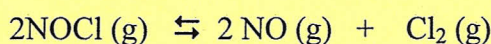
$$[A] = .0300 - .0054 = .0246 \text{ M}$$

$$[B] = .0180 - .0054 = .0126 \text{ M}$$

$$[AB] = .150 + .0054 = .1554 \text{ M}$$

$$\begin{aligned}
 \text{Solve Guess} &= .01 \\
 x &= .0054
 \end{aligned}$$

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.



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

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

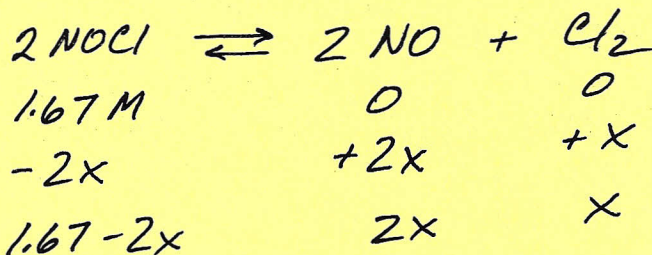
$$K = 0.0353$$

28% decomposes or breaks down

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

$$2x = .467$$

$$x = .233$$



$$K = \frac{[2x]^2 [x]}{[1.67 - 2x]^2}$$

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