The equilibrium concentration for the above reactions is:

$[H_2]_{eq} = 0.66$ M $\qquad [CO]_{eq} = 1.37$ M

The equilibrium expression is $Kc = \frac{[CH_3OH]^1}{[H_2]^2[CO]^1}$

$$K_c = \frac{0.200}{(0.66)^2 * 1.37} \qquad K_c = 0.335$$

If an equilibrium concentration is not given in the problem, but the equilibrium constant is given the above procedure is followed except X would be included in your expression and the quadratic is used to solve for X.

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

When using the quadratic to solve for the possibilities of X, one of the roots will result in a negative concentration at equilibrium. This root may be thrown out.

$[PCl_5] = \frac{.400}{2} = .200$ M or use moles they

c. The equilibrium constant, K_C, for the reaction below is 33.3 at 760 °C. If 0.400 mol of PCl_5 are placed in a 2.00 liter container, calculate the equilibrium concentrations of all species.

R	$PCl_5(g)$ <-->	$PCl_3(g)$ +	$Cl_2(g)$
I	.200 M	0	0
C	− X	+ X	+ X
E	.200 − X	+X	+ X

so $[PCl_5]_{eq} = .200 - .199 = .001$ M

$[PCl_3] = [Cl_2] = .199$ M

$$K_c = 33.3 = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{x^2}{.200 - x}$$

$33.3(.200 - x) = x^2$

$0 = x^2 + 33.3x - 6.66$

solve using quadratic $\quad x = -33.4988$ or $x = .199$ M $\quad [PCl_3]=0.199, [PCl_5]=0.001, [Cl_2]=0.199$

d. The equilibrium constant, K_C, for the reaction

R $\quad PCl_5(g)$ <--> $PCl_3(g) + Cl_2(g)$

is 33.3 at 760 °C. If 0.400 mol of PCl_5 and 1.0 mol of Cl_2 are placed in a 2.00 liter container, calculate the equilibrium concentrations of all species.

I	0.200 M	0	0.500 M
C	− X	+X	+ X
E	0.200 − X	X	0.500 + X

$$K_c = 33.3 = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{x(0.500 + x)}{(0.200 - x)}$$

$33.3(0.200 - x) = 0.500x + x^2$

$6.66 - 33.3x = 0.500x + x^2$

$0 = x^2 + 0.500x + 33.3x - 6.66$

$0 = x^2 + 33.8x - 6.66$

$x = 0.196$ M or -33.995 M

$[PCl_3]=0.196, [PCl_5]=0.004, [Cl_2]=0.696$

5. Kc = 19.9 for the reaction: R Cl_2 (g) + F_2 (g) ↔ 2ClF (g)

I 0.4 0.2 7.3

Find the equilibrium concentrations of all species if the initial concentration of [Cl_2] is 0.4M, [F_2] is 0.2 M, and [ClF] is 7.3 M.

C −x −x +2x

E 0.4−x 0.2−x 7.3+2x

$$19.9 = \frac{[ClF]^2}{[Cl_2][F_2]} = \frac{(7.3+2x)^2}{(0.4-x)(0.2-x)} = \frac{53.29 + 29.2x + 4x^2}{.08 - 0.6x + x^2}$$

$$19.9(.08 - 0.6x + x^2) = 53.29 + 29.2x + 4x^2$$

$$1.592 - 11.94x + 19.9x^2 = 53.29 + 29.2x + 4x^2$$

$$15.9x^2 - 41.14x - 51.698 = 0$$

$x = 3.5129$ or $-.9256$ ← choose this

$[Cl_2]_{eq} = 0.4 - (-.9256) = 1.3$ M

$[F_2]_{eq} = 1.1$ M

$[ClF]_{eq} = 7.3 + 2(-.9256) = 5.4$ M

6. Kc = 170 at 298K for the following reaction: $2NO_2$ (g) ↔ N_2O_4 (g)

a) Suppose the initial concentration of NO_2 is 0.015 M and the concentration of N_2O_4 is 0.025 M. Determine the equilibrium concentrations of all species. x is negligible, 170 >> .015

R $2NO_2(g)$ ⇌ $N_2O_4(g)$

I 0.015 M 0.025 M

C −2x +x

E 0.015−2x 0.025+x

$$Kc = 170 = \frac{[N_2O_4]}{[NO_2]^2} = \frac{.025+x}{(.015-2x)^2}$$

$$170(.015-2x)^2 = .025+x$$

$$170(0.000225 - 0.0006x + 4x^2) = .025 + x$$

$$.03825 - .102x + 680x^2 = .025$$

$$680x^2 - 0.102x + .01325 = 0$$

Complex roots - ignore the "i"

$x = .000075$ M

$[NO_2] = 0.015 - 2(.000075) \cong .015$ M

$[N_2O_4] = .025 + .000075 \cong .025$ M

7. Kc for the decomposition of ammonium hydrogen sulfide is 1.8 x 10^{-4} at 298K.

R NH_4HS (s) ↔ NH_3(g) + H_2S (g)

a) When the pure salt decomposes in a flask, what are the equilibrium concentration of NH_3 and H_2S?

I / 0 0

C +x +x

E x x

$$Kc = 1.8 \times 10^{-4} = [NH_3][H_2S] = x^2$$

$$\sqrt{1.8 \times 10^{-4}} = x = .013 \text{ M} = [NH_3]_{eq} = [H_2S]_{eq}$$

b) If NH_4HS is placed in a flask already containing 0.020 mol/L of NH_3 and then the system is allowed to come to equilibrium, what are the equilibrium concentrations of NH_3 and H_2S?

R $NH_4HS(s)$ ⇌ $NH_3(g)$ + $H_2S(g)$

I .020 M 0

C +x +x

E .020+x x

$$Kc = 1.8 \times 10^{-4} = [NH_3][H_2S] = (.020+x)(x)$$

$$1.8 \times 10^{-4} = .020x + x^2$$

$$0 = x^2 + .020x - 1.8 \times 10^{-4}$$

$x = -.02673$ M or 0.00673 M

$[H_2S]_{eq} = [NH_3]_{eq} = .0067$ M