Equil Notes: Practice Problems (Solutions)

1. a. Concentration of products + reactants are constant
   b. products
   c. solids & liquids

2. a. $K = \frac{[NO_2]^2}{[NO]^2[O_2]}$ b. $K = \frac{[Zn^{2+}]}{[Ag^+]^2}$ c. $K = \frac{[OH^-]^2}{[CO_3^{2-}]}$

3. a. $K>1$ product fav $K<0$ reactant fav $K=1$ neither favored
   b. Coefficient are exponents
   c. $\Delta n = (n_{products} - n_{reactants})$ *gases only!
   d. $M = \frac{mol}{L} \Rightarrow mol = M \times L$

4. a. $K = \frac{[C]^2}{[A]^2[B]} = \frac{[0.43]^2}{[0.55]^2[0.33]} = 1.9$ $K>1$ product favored

   (assume std cond & no gases)

   $K_p = K_c(RT)^{\Delta n} = 1.9(0.0821 \cdot 298)^0 = 1.9$ same

   b. $K = \frac{[CO][H_2O]}{[CO_2][H_2]} = \frac{(0.0046)^2}{(0.0954)(0.0454)} = 0.0049$

   $K<1$ reactant favored

   get M first

   $[CO] = [H_2O] = \frac{0.0092 mol}{2L} = 0.0046M$

   $[CO_2] = \frac{0.1908 mol}{2L} = 0.0954M$

   $[H_2] = \frac{0.0908 mol}{2L} = 0.0454M$

   $\Delta n = 0$

   $K_c = K_p = 0.0049$

5.a. nothing, K does not say anything about how fast (rate) a reaction goes only the conditions at equilibrium

b. Yes greatly K very large so highly product favored.

c. $K_P = K_C(RT)^{\Delta n} = 7.3\times10^{34}(0.0821\cdot298)^{(3-2)} = 1.8\times10^{36}$

d. $K_P = \frac{(P_{N_2})^2(P_{O_2})}{(P_{N_2O})^2} \Rightarrow x^2 \cdot 1.8\times10^{36} = \frac{(3.4\times10^{8})^2(2.4\times10^{7})}{x^2} \cdot x^2$

$P_{N_2O} = 1.2\times10^{-6}$ atm

$\frac{1.8\times10^{36}x^2}{1.8\times10^{36}} = \frac{2.8\times10^{24}}{1.8\times10^{36}}$

$\sqrt{x^2} = \sqrt{1.6\times10^{-12}}$

$x = 1.2\times10^{-6}$

e. Yes, P_{N_2O} should be a very small amount since the reaction is so much product favored there should be very little reactant @ equilibrium.

6. a. K is cubed (K^3)

b. Take the reciprical ($\frac{1}{K}$)

c. $K_{overall} = K_1 \times K_2 \times \ldots$ multiply the K's of each step

7. Original eqn + K $2N_2O \rightleftharpoons 2N_2 + O_2$ $K = 7.3 \times 10^{3}$

a. Coeff of rxn are doubled so new $K = K^2$

$K = 5.3 \times 10^{69}$

b. Rxn is flipped so new $K = \frac{1}{K}$

$K = 1.4 \times 10^{-35}$

8. Reaction Quotient (Q) - may or may not be @ equil

Equilibrium Constant (K) - @ equil

9. $Q = \frac{[HI]^2}{[H_2][I_2]} = \frac{(3.21 \times 10^{3})^2}{(3.21 \times 10^{-3})(3.21 \times 10^{-3})} = 1$

$12.3 \quad 1$

$K > Q$

→ reaction proceeds to the right since Q is less needs to make more products to increase to reach K

10. a. The question gives/asks information about both initial & equilibrium conditions

b. Calculate Q first then compare Q to K

$K > Q$ → $K < Q$ ← $K = Q$ @ equil

c. If you end up w/ an expression of $ax^2+bx+c=0$

d. If K is very small ie $\times 10^{-5}$ magnitude

11. a. R $S_{2(g)} \rightleftharpoons 2S_{(g)}$

I 0.008 (Y)

C $-x$ $+2x$

E $0.008-x$ 4.0×10^{-11}

$[S_2] = 0.008 - 2.0\times10^{-11} = 0.008$ ← x is very small so $[S_2]$ doesn't change

$K = \frac{[S]^2}{[S_2]} = \frac{(0.008)^2}{(2.0\times10^{-11})} = 4.0\times10^{8}$

$[S_2]_{initial} = \frac{0.0040\,mol}{0.500\,L} = 0.008\,M$

$[S]_{equil} = \frac{2.0\times10^{-11}\,mol}{0.500\,L} = 4.0\times10^{-11}\,M$

$\frac{2x}{2} = \frac{4.0\times10^{-11}}{2}$

$x = 2.0\times10^{-11}$

11. b.

R	N_2O_4	$\rightleftharpoons$	$2NO_2$
I	0.67		0
C	$-x$		$+2x$
E	0.56		$2x$

$[N_2O_4]_{initial} = \frac{0.50\,mol}{0.750\,L} = 0.67M$

$[N_2O_4]_{equil} = \frac{0.42\,mol}{0.750\,L} = 0.56M$

$0.67 - x = 0.56$

$x = 0.11$

$[NO_2] = 2x = 2(0.11) = 0.22M$

$K = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(0.22)^2}{(0.56)} = 0.086$

c.

R	I_2	+	Cl_2	$\rightleftarrows$	$2ICl$
I	0.100		0.100		0.100
C	$-x$		$-x$		$+x$
E	$0.100-x$		$0.100-x$		$0.100+x$

Check Q first ← since you have amounts for each

$Q = \frac{[ICl]^2}{[I_2][Cl_2]} = \frac{0.100^2}{0.100^2} = 1$

$K > Q$

$81.9 \rightarrow 1$ forward

$K = \frac{[ICl]^2}{[I_2][Cl_2]} \Rightarrow \sqrt{81.9} = \sqrt{\frac{(0.100+x)^2}{(0.100-x)^2}}$

$9.05 = \frac{0.100+x}{0.100-x}$ (multiply by $(0.100-x)$)

$0.905 - 9.05x = 0.100 + x$

$-0.100 + 9.05x \quad -0.100 + 9.05x$

$\frac{0.805}{10.05} = \frac{10.05x}{10.05}$

$0.0801 = x$

$[I_2] = 0.100 - x = 0.020M$ ($x = 0.0801$)

$[Cl_2]$

$[ICl] = 0.100 + x = 0.180M$ ($x = 0.0801$)

11.d.

	SO_2	+ NO_2	$\rightleftharpoons$	NO	+ SO_3
R			←		
I	0.0025	0.0025		0.025	0.025
C	+x	+x		−x	−x
E	0.0025+x	0.0025+x		0.025−x	0.025−x

Amounts for each so check Q 1st

$$Q = \frac{[NO][SO_3]}{[SO_2][NO_2]} = 100$$

$K < Q$ rxn goes left

85.0 100

$$K = \frac{[NO][SO_3]}{[SO_2][NO_2]}$$

$$\sqrt{85.0} = \sqrt{\frac{(0.025-x)^2}{(0.0025+x)^2}}$$

$$(0.0025+x)\,9.22 = \frac{0.025-x}{0.0025+x}(0.0025+x)$$

$$0.0230 + 9.22x = 0.025 - x$$

$$-0.0230 \quad -0.0230 + x$$

$$\frac{10.22x}{10.22} = \frac{0.002}{10.22}$$

$$x = 0.000196$$

$[SO_2] + [NO_2] = 0.0025 + 0.000196 = 0.0027$ M

$[NO]$ & $[SO_2] = 0.025 - 0.000196 = 0.025$ M

11e.

	PCl_5	$\rightleftharpoons$	PCl_3	+	Cl_2
I	3.50		0.50		0.50
C	+x		−x		−x
E	3.50+x		0.50−x		0.50−x

$Q = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{(0.50)^2}{3.50} = 0.0714$

$K < Q$

0.0454 0.0714

goes to the left

CANNOT $\sqrt{\ }$!

$K = \frac{[PCl_3][Cl_2]}{[PCl_5]}$ $0.0454 = \frac{(0.50-x)^2}{3.50+x}$ $(3.50+x)$

$0.159 + 0.0454x = (0.50-x)(0.50-x)$ FOIL

$0.159 + 0.0454x = 0.25 - x + x^2$ get = 0

$-0.159 \quad -0.0454x \quad -0.159 - 0.0454x$

$0 = 0.091 - 1.0454x + x^2$

a b c

must use quadratic formula $x = \frac{-b \pm \sqrt{b^2-4ac}}{2}$

You get two x's $(x = 0.10)$ pick the x that gives you a positive [ ] for equil b/c a neg [ ] isn't possible

$[PCl_3] = 0.50 - 0.10 = 0.40\ M$

$[Cl_2]$

$[PCl_5] = 3.50 + 0.10 = 3.60\ M$

11. f. R $2\,NH_3 \rightleftharpoons N_2 + 3H_2$

I	$[NH_3]$	0	0	
C	$-2x$	$+x$	$+3x$	
E	$[NH_3]-2x$	x	0.342 ← given	$\frac{3x}{3} = \frac{0.342}{3}$
	$[NH_3]-2(0.114)$	0.114		$x = 0.114$

$$K = \frac{[N_2][H_2]^3}{[NH_3]^2}$$

$$6.30 = \frac{(0.114)(0.342)^3}{([NH_3]-0.228)^2}$$

↳ change to y to make it easier to solve

$$6.30 = \frac{0.00456}{(y-0.228)^2}$$

$(y-0.228)^2$ ↑ get this on top

$$\frac{(y-0.228)^2 6.30}{6.30} = \frac{0.00456}{6.30}$$

$$\sqrt{(y-0.228)^2} = \sqrt{0.000724}$$

$$y - 0.228 = 0.0269$$

$$+0.228 \quad +0.228$$

$$y = 0.255$$

$$[NH_3] = 0.255M$$