AP Equilibrium Review for the Final Exam!!!

Key

$4\,KO_2(s) + 2\,H_2O(g)$ <-----> $4\,KOH(s) + 3\,O_2(g)$

$K_c = \frac{[O_2]^3}{[H_2O]^2}$ $K_p = \frac{P_{O_2}^{3}}{P_{H_2O}^{2}}$

$P_4(s) + 5\,O_2(g)$ <-----> $P_4O_{10}(s)$

$K_c = \frac{1}{[O_2]^5}$ $K_p = \frac{1}{(P_{O_2})^5}$

1. a. Write an expression for Kc and for Kp for each equation, above.

b. What quantities and units should be used in Kc calcualtions? in Kp calculations?

Kc: plug in concentration in moles / Liters (molarity!)

Kp: plug in partial pressures in atmospheres.

2. Write an expression for Kc for each equation:

a. $H_2CO_{3(aq)} + 2\,H_2O_{(l)}$ ------> $2\,H_3O^{+1}{}_{(aq)} + CO_3{}^{-2}{}_{(aq)}$

$\boxed{K_c = \frac{[H_3O^{+1}]^2[CO_3^{-2}]}{[H_2CO_3]}}$

b. $Cu_3(PO_4)_{2\,(s)}$ <-----> $3\,Cu^{+2}{}_{(aq)} + 2\,PO_4{}^{-3}{}_{(aq)}$

$\boxed{K_c = [Cu^{+2}]^3[PO_4^{-3}]^2}$

The Kc for (b) involves a specific type of Kc called a $\boxed{K_{sp}}$

("solubility product")

3. $2\,NOCl_{(g)}$ <-----> $2\,NO_{(g)} + Cl_{2(g)}$ $K_p = 1.6 \times 10^{-5}$ at 35°C.

(assume 35°C for this whole problem).

a. Find Kp of this rxn: $2\,NO_{(g)} + Cl_{2(g)}$ <-----> $2\,NOCl_{(g)}$

$K_p = \left(\frac{1}{1.6 \times 10^{-5}}\right) = 62500 \rightarrow \boxed{63000 \text{ or } 6.3 \times 10^4}$

b. Find Kp of this rxn: $4\,NOCl_{(g)}$ <-----> $4\,NO_{(g)} + 2\,Cl_{2(g)}$

$K_p = (1.6 \times 10^{-5})^2 = \boxed{2.6 \times 10^{-10}}$

c. Find Kp of this rxn: $NO_{(g)} + \frac{1}{2}\,Cl_{2(g)}$ <-----> $NOCl_{(g)}$

$K_p = \left(\frac{1}{1.6 \times 10^{-5}}\right)^{1/2} = \boxed{250}$

d. Find Kc of this rxn: $2\,NOCl_{(g)}$ <-----> $2\,NO_{(g)} + Cl_{2(g)}$

$K_p = K_c(RT)^{\Delta n}$

$1.6 \times 10^{-5} = K_c(.0821 \cdot 308)^1$

$\boxed{K_c = 6.3 \times 10^{-7}}$

4. $CH_6N_2O_{2(s)}$ <-----> $2\ NH_{3(g)} + CO_{2(g)}$ Kp = 0.0029 at 25°C.

a. Calculate the partial pressure of each gas, and the total gas pressure, at equilibrium.

b. Calculate Kc of this rxn at 25°C.

(a)

$$CH_6N_2O_{2(s)} \rightleftharpoons 2NH_{3(g)} + CO_{2(g)}$$

$$\quad\quad\quad\quad 2x \quad\quad x$$

$$K_p = (p_{NH_3})^2(p_{CO_2}) = (2x)^2 \cdot x = .0029$$

$$4x^3 = .0029$$

$$x = .08984$$

(a)

$$p_{NH_3} = 2x = 0.18\text{ atm}$$

$$p_{CO_2} = x = 0.090\text{ atm}$$

$$p_{TOTAL} = 3x = 0.27\text{ atm}$$

$$K_p = K_c(RT)^{\Delta n}$$

$$.0029 = K_c(.0821 \cdot 298)^{+3}$$

$$K_c = 1.98\times10^{-7} \rightarrow \boxed{K_c = 2.0\times10^{-7}}$$ (b)

5. $2\ HI_{(g)}$ <-----> $H_{2(g)} + I_{2(g)}$ Kp = 0.0218

If 3.00 atm of HI are placed into an evacuated flask, and this reaction reaches equilibrium, what will be the partial pressures of each gas at equilibrium?

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

	$2HI$	H_2	I_2
Initial	3.00	0	0
C	$-2x$	$+x$	$+x$
E	$3.00-2x$	x	x

$$K_p = \frac{(p_{H_2})(p_{I_2})}{(p(HI))^2} = \frac{x^2}{(3.00-2x)^2} = 0.0218$$

← (you could just square root each side and save some time...)

$$x^2 = .0218(3.00-2x)^2 = .0218(9.00 \overbrace{-6x-6x}^{-12x} + 4x^2)$$

$$x^2 = 0.1962 - 0.2616x + .0872x^2$$

$$0.9128x^2 + .2616x - .1962 = 0$$

$$x = \frac{-0.2616 \pm \sqrt{(0.2616)^2 - 4(.9128)(-.1962)}}{2(.9128)}$$

$$x = \frac{-.2616 \pm .88589}{1.8256} = 0.34196 \text{ or } \cancel{-0.62855}$$

$$p_{H_2} = p_{I_2} = 0.34196 \rightarrow \boxed{0.342\text{ atm}}$$

$$P_{HI} = 3.00 - 2x = \boxed{2.32\text{ atm}}$$

6. $2\,HI_{(g)}$ <-----> $H_{2(g)} + I_{2(g)}$ Kc = 0.0423

If 3.00 moles of HI are placed into an evacuated flask with a volume of 5.00 liters, and this reaction reaches equilibrium, what will be the molarities of each gas present at equilibrium?

$$\frac{3.00\text{ moles}}{5.00\text{ L}} = 0.600\text{ M}$$

$$\boxed{\begin{array}{l}[H_2] = [I_2] = .0874\text{ M} \\ [HI] = .600-2x = 0.425\text{ M}\end{array}}$$

	$2HI_{(g)}$ $\rightleftarrows$	$H_{2(g)}$ +	$I_{2(g)}$
I	0.600 M	0	0
C	$-2x$	$+x$	$+x$
E	$.600-2x$	x	x

(Since you have "x" in the chart) you need to do the chart in molarities (not moles). Since $\Delta n_{gas} = 0$ you'd get the right answer if you do moles, and change to [ ] at end, but only since $\Delta n_{gas} = 0$!

$$K_c = \frac{[H_2][I_2]}{[HI]^2} = \frac{x^2}{(.600-2x)^2} = 0.0423$$

$$x^2 = .0423(.600-2x)^2 = .0423(0.36 - 2.4x + 4x^2)$$

$$x^2 = 0.015228 - .10152x + 0.1692x^2$$

$$0.8308x^2 + .10152x - .015228 = 0$$

$$x = \frac{-.10152 \pm \sqrt{(.10152)^2 - 4(.8308)(-.015228)}}{2(.8308)} \rightarrow x = \frac{-.10152 \pm .24680}{1.6616}$$

$x = .08743$ or $\cancel{-.2096}$

7. $2\,NO_{(g)} + Br_{2(g)}$ <-----> $2\,NOBr_{(g)}$

0.512 moles NO, 0.502 moles Br_2, and 0.096 moles of NOBr are placed into a rigid 12.0 liter flask. Once the mixture reaches equilibrium, 0.118 moles of NOBr are present.

Calculate Kc.

	$2NO_{(g)}$ +	$Br_{2(g)}$ $\rightleftarrows$	$2\,NOBr_{(g)}$
I	.512	.502	.096
C	$-.022$	$-.011$	$+.022$
E	0.490	0.491	.118

Since no "x" values in chart, can do moles or molarities in I.C.E. chart but get molarities before plugging into K_c.

$$[NO] = \frac{.490\text{ mole}}{12.0\text{ L}} = .04083\text{ M}$$

$$[Br_2] = \frac{.491\text{ mole}}{12.0\text{ L}} = 0.04092\text{ M}$$

$$[NOBr] = 0.118\text{ mole} / 12.0\text{ L} = 0.00983\text{ M}$$

$$K_c = \frac{[NOBr]^2}{[NO]^2[Br_2]} = \frac{(.00983)^2}{(.04083)^2(.04092)} = \boxed{1.42}$$

8. $2\,NO_{(g)} + Br_{2(g)} \longleftrightarrow 2\,NOBr_{(g)}$

The above 3 gases are placed into a rigid flask at the following partial pressures:
pNO = 0.1163 atm, pBr2 = 0.0478 atm, and pNOBr = 0.0132 atm.
Once equilibrium is established, NO is present at a partial pressure of 0.0526 atm.
Calculate Kp.

	$2NO_{(g)}$ +	$Br_{2(g)} \rightleftarrows$	$2NOBr_{(g)}$
I	.1163	.0478	.0132
C	−.0637	−.03185 *	+.0637
E	.0526	0.01595	.0769

$* .0637\left(\frac{1}{2}\right) = .03185$

$$K_p = \frac{(p_{NOBr})^2}{(p_{NO})^2(p_{Br_2})} = \frac{(.0769)^2}{(.0526)^2(.01595)} = \boxed{134}$$

9. PbI_2 has a Ksp of 1.4×10^{-8}.

a. Find the molar solubility of lead iodide into water, and also calculate the concentrations of lead (II) and iodide ion in solution at equilibrium.

$PbI_{2(s)} \rightleftarrows Pb^{+2}_{(aq)} + 2I^{-1}_{(aq)}$

x   2x

$K_{sp} = [Pb^{+2}][I^{-1}]^2$

$1.4\times10^{-8} = x\,(2x)^2$

$4x^3 = 1.4\times10^{-8}$

$x = .001518\ M$

(.0015 moles of PbI_2 can dissolve per liter of solution)

$\boxed{[Pb^{+2}] = .0015\ M \quad [I^-] = .0030\ M \quad \text{molar solubility} = .0015\ M}$

b. Find the molar solubility of lead iodide into a solution of 0.20 M KI, and also calculate the concentrations of lead (II) and iodide ion in solution at equilibrium.

	$PbI_{2(s)} \rightleftarrows$ $Pb^{+2}_{(aq)}$ +	$2I^{-1}_{(aq)}$
I	0	.20
C	+x	+2x
E	x	.20+2x

$K_{sp} = [Pb^{+2}][I^-]^2 = x(.20+2x)^2 = 1.4\times10^{-8}$

assume $2x \ll .20$

$x(.20)^2 = 1.4\times10^{-8}$

$x = 3.5\times10^{-7}\ M$

$\boxed{\text{molar solubility} = 3.5\times10^{-7}\ M}$

(3.5×10^{-7} moles PbI_2 can dissolve per liter of solution)

$\boxed{[Pb^{+2}] = 3.5\times10^{-7}\ M \quad [I^-] = 0.20\ M}$

$.20 + 2x \simeq 0.20!$

9c. 100. mL of 0.40 M $Pb(NO_3)_2$ are mixed with 80. mL of 0.60 M KI. (K_{sp} of $PbI_2 = 1.4\times10^{-8}$)
Calculate the concentrations of of lead II and iodide ions once the reaction has occurred, and equilibrium has been established.

$(.100\ L)(.40\ \frac{mole}{L}) = 0.040$ moles $Pb(NO_3)_2$

$(.080L)(.60\ \frac{mole}{L}) = 0.048$ moles KI

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \longrightarrow PbI_2 + 2KNO_3$$

	$Pb(NO_3)_2$	$2KI$
I	.040	.048
C	−.024	−.048
F	.016 moles	0

$$[Pb^{+2}] = \frac{0.016\text{ moles}}{0.180\ L} = 0.08888 \rightarrow \boxed{0.089\ M\ \ Pb^{+2}}$$

100. mL
+80. mL
180. mL
total V

Now do Ksp to find out [I-].

$$PbI_{2(s)} \rightleftarrows Pb^{+2}_{(aq)} + 2\,I^{-1}_{(aq)}$$

	Pb^{+2}	I^{-1}
I	$.088\bar{8}$	0
C	$+\frac{x}{2}$	+x
E	$.0888+\frac{x}{2}$	x

or +x, +2x and [I-] = 2x at end

$$K_{sp} = [Pb^{+2}][I^-]^2 = (.0888+\tfrac{x}{2})\,x^2 = 1.4\times10^{-8}$$

assume $\frac{x}{2} \ll .0888$

$(.0888)\,x^2 = 1.4\times10^{-8}$

$x = .0003969$ M

$$\boxed{[I^-] = x = .00040\ M}$$

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~~9~~. Ag_3PO_4 has a solubility of 0.00196 grams per liter.

a. Convert this to a molar solubility.

$$\left(\frac{.00196\ g}{L}\right)\left(\frac{1\ mole}{418.575\ g}\right) = 4.6825\times10^{-6}\ M \rightarrow \boxed{4.68\times10^{-6}\ M}$$

b. Calculate Ksp for this compound.

$$Ag_3PO_{4(s)} \rightleftarrows 3\,Ag^{+1}_{(aq)} + PO_4^{-3}$$

$3(4.6825\times10^{-6})$ or 1.40476×10^{-5} ; 4.6825×10^{-6}

$$K_{sp} = [Ag^{+1}]^3[PO_4^{-3}]$$
$$= (1.40476\times10^{-5})^3(4.6825\times10^{-6})$$
$$= 1.298\times10^{-20} \rightarrow \boxed{1.30\times10^{-20}}$$

c. If 200. mL of 0.300 M $AgNO_3$ are mixed with 100. mL of 0.400 M Na_3PO_4, calculate the concentration of silver ion and phosphate ion after the reaction has occurred and equilibrium has been established. Ag_3PO_4 has a Ksp of 1.3×10^{-20}.

$$(.200\ L)\left(.300\ \frac{mol}{L}\right) = .0600 \text{ moles } AgNO_3$$
$$(.100\ L)\left(.400\ \frac{mol}{L}\right) = .0400 \text{ moles } Na_3PO_4$$

$$3\,AgNO_{3(aq)} + Na_3PO_{4(aq)} \longrightarrow Ag_3PO_{4(s)} + 3\,NaNO_{3(aq)}$$

	$3\,AgNO_{3(aq)}$	$Na_3PO_{4(aq)}$
I	.0600	.0400
C	−.0600	−.0200
F	0	.0200

$$[PO_4^{-3}] = \frac{0.0200\ moles}{0.300\ L} = 0.0666\overline{6} \rightarrow \boxed{0.0667\ M}$$

0.0667 M

$$K_{sp} = [Ag^{+1}]^3[PO_4^{-3}]$$
$$1.3\times10^{-20} = (3x)^3(.06667 + x)$$

$Ag_3PO_{4(s)} \rightleftarrows$	$3\,Ag^{+1}_{(aq)}$	$+\ PO^{-3}_{4(aq)}$
	0	.06667
	+3x	+x
	3x	.06667 + x

or do x and $\frac{x}{3}$

assume ~~.066~~ $x \ll .06667$

$$1.3\times10^{-20} = (3x)^3(.06667)$$

$\cancel{x = 8.972\times10^{-8}\ M}$ $x = 1.93\times10^{-7}\ M$

$[Ag^{+1}] = 3x =$ $\cancel{2.6917\times10^{-7}\ M}$ $5.7998\times10^{-7}\ M$

$$\boxed{[Ag^{+}] = 5.8\times10^{-7}\ M}$$