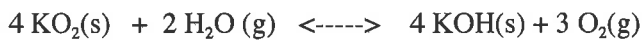


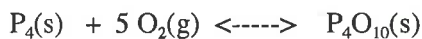
AP Equilibrium Review for the Final Exam!!!

Key



$$K_c = \frac{[\text{O}_2]^3}{[\text{H}_2\text{O}]^2}$$

$$K_p = \frac{P_{\text{O}_2}^3}{P_{\text{H}_2\text{O}}^2}$$



$$K_c = \frac{1}{[\text{O}_2]^5}$$

$$K_p = \frac{1}{(P_{\text{O}_2})^5}$$

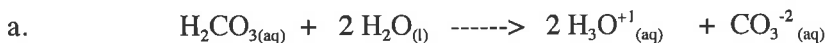
1. a. Write an expression for K_c and for K_p for each equation, above.

b. What quantities and units should be used in K_c calculations? in K_p calculations?

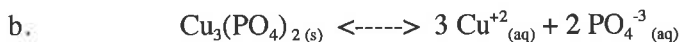
K_c : plug in concentration in moles / Liters (molarity!)

K_p : plug in partial pressures in atmospheres.

2. Write an expression for K_c for each equation:



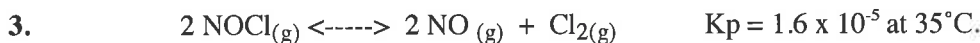
$$K_c = \frac{[\text{H}_3\text{O}^+]^2 [\text{CO}_3^{2-}]}{[\text{H}_2\text{CO}_3]}$$



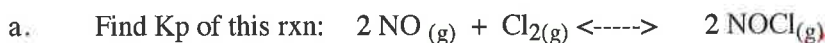
$$K_c = [\text{Cu}^{2+}]^3 [\text{PO}_4^{3-}]^2$$

The K_c for (b) involves a specific type of K_c called a K_{sp}

("solubility product")



(assume 35°C for this whole problem).

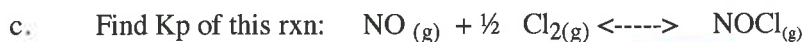


$$K_p = \left(\frac{1}{1.6 \times 10^{-5}} \right) = 62500 \rightarrow$$

$$63000 \text{ or } 6.3 \times 10^4$$



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



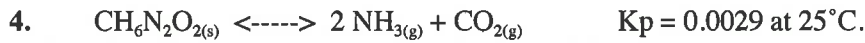
$$K_p = \left(\frac{1}{1.6 \times 10^{-5}} \right)^{1/2} = 250$$



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

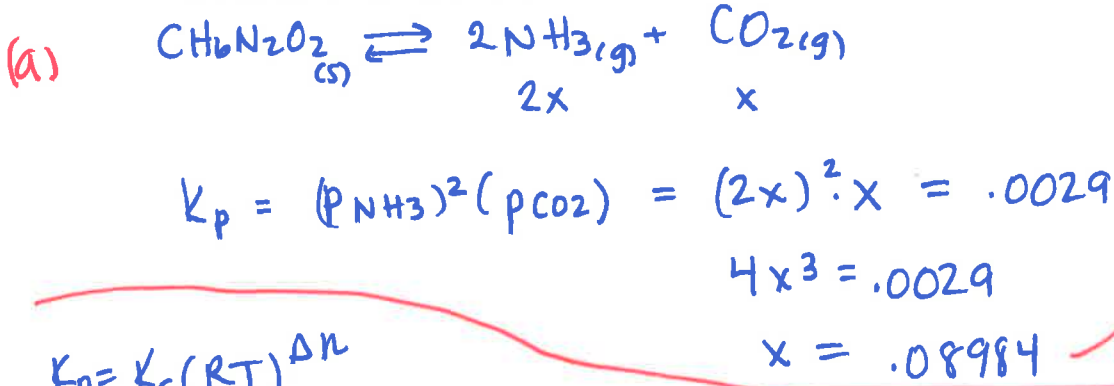
$$1.6 \times 10^{-5} = K_c (.0821 \cdot 308)^1$$

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



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

b. Calculate K_c of this rxn at 25°C .



(a)

$$P_{\text{NH}_3} = 2x = 0.18 \text{ atm}$$

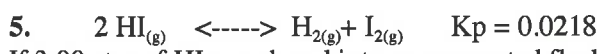
$$P_{\text{CO}_2} = x = 0.090 \text{ atm}$$

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

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

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

$$K_c = 1.98 \times 10^{-7} \rightarrow K_c = 2.0 \times 10^{-7} \text{ (b)}$$



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?

	$2\text{HI}(\text{g})$	\rightleftharpoons	$\text{H}_2(\text{g})$	$+$	$\text{I}_2(\text{g})$
Initial	3.00		0		0
C	$-2x$		$+x$		$+x$
E	$3.00 - 2x$		x		x

$$K_p = \frac{(P_{\text{H}_2})(P_{\text{I}_2})}{(P_{\text{HI}})^2} = \frac{x^2}{(3.00 - 2x)^2} = 0.0218$$

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

$$x^2 = .0218 (3.00 - 2x)^2 = .0218 (9.00 - 12x + 4x^2)$$

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

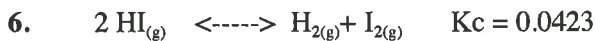
$$0.9128x^2 + .2616x - .1962 = 0$$

$$x = \frac{-0.2616 \pm \sqrt{(0.2616)^2 - 4(0.9128)(-0.1962)}}{2(0.9128)}$$

$$x = \frac{-0.2616 \pm .88589}{1.8256} = 0.34196 \text{ or } -0.62855$$

$$P_{\text{H}_2} = P_{\text{I}_2} = 0.34196 \rightarrow 0.342 \text{ atm}$$

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



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?

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

	$2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$		
I	0.600 M	0	0
C	$\frac{-2x}{}$	$\frac{+x}{}$	$\frac{+x}{}$
E	$.600 - 2x$	x	x

$[\text{H}_2] = [\text{I}_2] = .0874 \text{ M}$
 $[\text{HI}] = .600 - 2x = 0.425 \text{ M}$

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

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{x^2}{(.600 - 2x)^2} = 0.0423$$

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

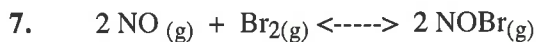
$$x^2 = 0.015228 - .10152x + 0.1692x^2$$

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

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

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

$$x = \frac{-.10152 \pm .24680}{1.6616}$$



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 K_c .

	$2 \text{NO}_{(g)} + \text{Br}_{2(g)} \rightleftharpoons 2 \text{NOBr}_{(g)}$		
I	.512	.502	.096
C	$\frac{-.022}{}$	$\frac{-.011}{}$	$\frac{+.022}{}$
E	.490	.491	.118

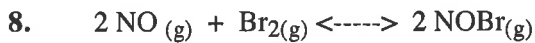
Since no "x" values in chart, can do moles or molarities in I.C.E. chart but get molarities before plugging into K_c .

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

$$[\text{Br}_2] = \frac{.491 \text{ mole}}{12.0 \text{ L}} = 0.04092 \text{ M}$$

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

$$K_c = \frac{[\text{NOBr}]^2}{[\text{NO}]^2[\text{Br}_2]} = \frac{(0.00983)^2}{(.04083)^2(.04092)} = 1.42$$



The above 3 gases are placed into a rigid flask at the following partial pressures:

$p_{\text{NO}} = 0.1163 \text{ atm}$, $p_{\text{Br}_2} = 0.0478 \text{ atm}$, and $p_{\text{NOBr}} = 0.0132 \text{ atm}$.

Once equilibrium is established, NO is present at a partial pressure of 0.0526 atm.

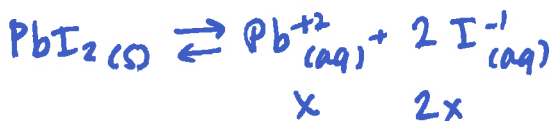
Calculate K_p .

	$2 \text{NO}_{(g)}$	$+$	$\text{Br}_{2(g)}$	\rightleftharpoons	$2 \text{NOBr}_{(g)}$	
I	.1163		.0478		.0132	
C	<u>-.0637</u>		<u>-.03185*</u>		<u>+.0637</u>	*.0637 ($\frac{1}{2}$) = .03185
E	.0526		0.01595		.0769	

$$K_p = \frac{(p_{\text{NOBr}})^2}{(p_{\text{NO}})^2 (p_{\text{Br}_2})} = \frac{(.0769)^2}{(.0526)^2 (.01595)} = \boxed{134}$$

9. PbI_2 has a K_{sp} 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.



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

$$1.4 \times 10^{-8} = x(2x)^2$$

$$4x^3 = 1.4 \times 10^{-8}$$

$$x = .001518 \text{ M}$$

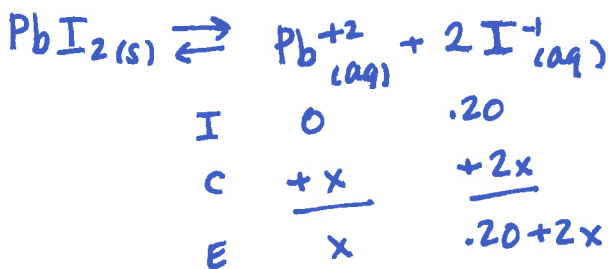
(.0015 moles of PbI_2 can dissolve per liter of solution)

$$[\text{Pb}^{+2}] = .0015 \text{ M}$$

$$[\text{I}^{-1}] = .0030 \text{ M}$$

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.



molar solubility = $3.5 \times 10^{-7} \text{ M}$

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

$$[\text{Pb}^{+2}] = 3.5 \times 10^{-7} \text{ M}$$

$$[\text{I}^{-1}] = 0.20 \text{ M}$$

$$K_{sp} = [\text{Pb}^{+2}][\text{I}^{-1}]^2 = x(.20+2x)^2 = 1.4 \times 10^{-8}$$

assume $2x \ll .20$

$$x(.20)^2 = 1.4 \times 10^{-8}$$

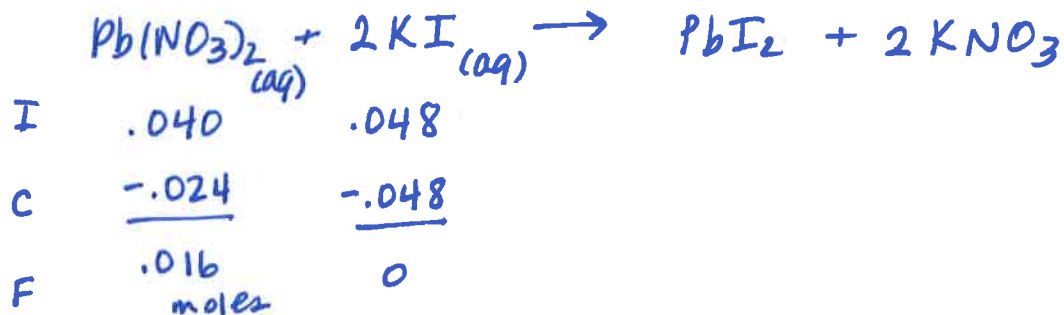
$$x = 3.5 \times 10^{-7} \text{ M}$$

↑
 $.20 + 2x \approx .20!$

9c 100. mL of 0.40 M $\text{Pb}(\text{NO}_3)_2$ are mixed with 80. mL of 0.60 M KI. (K_{sp} of $\text{PbI}_2 = 1.4 \times 10^{-8}$)
 Calculate the concentrations of lead II and iodide ions once the reaction has occurred, and equilibrium has been established.

$$(.100 \text{ L}) (.40 \frac{\text{mole}}{\text{L}}) = 0.040 \text{ moles } \text{Pb}(\text{NO}_3)_2$$

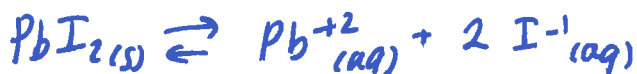
$$(.080 \text{ L}) (.60 \frac{\text{mole}}{\text{L}}) = 0.048 \text{ moles } \text{KI}$$



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

$$\begin{array}{l} \nearrow \\ 100. \text{ mL} \\ + 80. \text{ mL} \\ \hline 180. \text{ mL} \\ \text{total V} \end{array}$$

Now do K_{sp} to find out $[\text{I}^-]$.



I	.0888	0
C	<u>$+\frac{x}{2}$</u>	<u>$+x$</u>
E	$.0888 + \frac{x}{2}$	x

) or $+x$, $+2x$
 and $[\text{I}^-] = 2x$
 at end

$$K_{sp} = [\text{Pb}^{+2}][\text{I}^-]^2 = (.0888 + \frac{x}{2}) x^2 = 1.4 \times 10^{-8}$$

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

$$(.0888) x^2 = 1.4 \times 10^{-8}$$

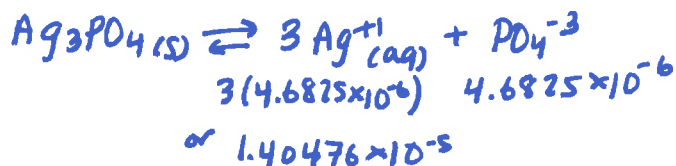
$$x = .0003969 \text{ M}$$

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

10. a. Ag_3PO_4 has a solubility of 0.00196 grams per liter.
 a. Convert this to a molar solubility.

$$\left(\frac{0.00196 \text{ g}}{\text{L}}\right) \left(\frac{1 \text{ mole}}{418.575 \text{ g}}\right) = 4.6825 \times 10^{-6} \text{ M} \rightarrow \boxed{4.68 \times 10^{-6} \text{ M}}$$

- b. Calculate Ksp for this compound.



$$K_{sp} = [\text{Ag}^+]^3 [\text{PO}_4^{3-}]$$

$$= (1.40476 \times 10^{-5})^3 (4.6825 \times 10^{-6})$$

$$= 1.298 \times 10^{-20} \rightarrow \boxed{1.30 \times 10^{-20}}$$

- 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} .

$$(0.200 \text{ L})(0.300 \frac{\text{mol}}{\text{L}}) = 0.0600 \text{ moles AgNO}_3$$

$$(0.100 \text{ L})(0.400 \frac{\text{mol}}{\text{L}}) = 0.0400 \text{ moles Na}_3\text{PO}_4$$



I	.0600	.0400
C	<u>-.0600</u>	<u>-.0200</u>
F	0	.0200

$$[\text{PO}_4^{3-}] = \frac{0.0200 \text{ moles}}{0.300 \text{ L}} = 0.066\bar{6} \rightarrow \boxed{0.0667 \text{ M}}$$

← 0.0667 M

$$K_{sp} = [\text{Ag}^+]^3 [\text{PO}_4^{3-}]$$

$$1.3 \times 10^{-20} = (3x)^3 (0.06667 + x)$$

assume ~~the~~
 $x \ll 0.06667$

$$1.3 \times 10^{-20} = (3x)^3 (0.06667)$$

$$x = \cancel{8.972 \times 10^{-8} \text{ M}} \quad x = 1.93 \times 10^{-7} \text{ M}$$

$$[\text{Ag}^+] = 3x = \cancel{2.6917 \times 10^{-7} \text{ M}} \quad 5.7998 \times 10^{-7} \text{ M}$$

$$\boxed{[\text{Ag}^+] = 5.8 \times 10^{-7} \text{ M}}$$



or do

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

0	.06667
<u>+3x</u>	<u>+x</u>
3x	.06667 + x