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**Unit#7**  
**AP Chem**  
**Topic#8**  
**Equilibrium**

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Student Edition

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Equilibrium  
Topic#8

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# General Equilibrium

**Equilibrium**  
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**Formula Page**

**EQUILIBRIUM**

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

$$K_w = [OH^-][H^+] = 1.0 \times 10^{-14} \text{ @ } 25^\circ\text{C}$$
$$= K_a \times K_b$$

$$pH = -\log[H^+], pOH = -\log[OH^-]$$
$$14 = pH + pOH$$

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pOH = pK_b + \log \frac{[HB^+]}{[B]}$$

$$pK_a = -\log K_a, pK_b = -\log K_b$$

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

where  $\Delta n$  = moles product gas – moles reactant gas

**Equilibrium Constants**

$K_a$  (weak acid)

$K_b$  (weak base)

$K_w$  (water)

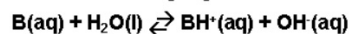
$K_p$  (partial pressures; gases)

$K_c$  (molar concentrations)

$K_{sp}$  (solubility product;  
insoluble ionic compounds)



$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$



$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

**Half Titration Point**

$$[A^-] = [HA]$$

$$pH = pK_a$$

or

$$[HB^+] = [B]$$

$$pOH = pK_b$$

**ICE Tables**

Elimination of  $\pm x$   
Ionization Rule  
( $< 5\%$  ionization)

**Electrochemistry**

$$K = 10^{(nE/0.0592)}$$

$$E^\circ_{\text{cell}} = \frac{0.0592}{n} \log K$$

**Thermochemistry**

$$K = e^{(-\Delta G/2.48)}$$

$$\Delta G = -RT \ln K$$

$$K_{\text{eq}} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

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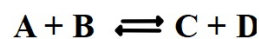
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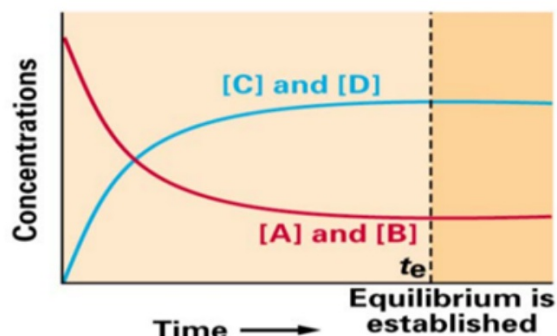
**Equilibrium Intro**

**Generalities of Equilibrium**

- State where the rate of the forward reaction is equal to the rate of the reverse reaction.
- Concentration of all species remains constant (constant  $T$ ).
- Reactions are reversible, yet some have great obstacles going in one direction.
  - High  $E_a$ , gas release, solid formed (precipitation), etc.
- Dynamic equilibrium - reaction going forward and in reverse.
- Take a look at the reaction of carbon monoxide and steam in a closed vessel:
  - Does the reaction favor reactants or products?
  - What can you say about the relative energies of reactants and products?
  - What can be you say about the reaction rate? Why?
  - What can you say about the degree of organization?



@ equilibrium  $k_f = k_r$



Is  $K_{eq} > 1$  or  $< 1$  for the above graph?

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**Equilibrium Constant/Expression**

**Significance of  $K$**

- $K > 1$       favors product formation at equilibrium ( $1.0 \times 10^4$ )
- $K \gg 1$      almost entirely products at equilibrium ( $1.0 \times 10^{23}$ )
- $K < 1$       favors reactants at equilibrium ( $1.0 \times 10^{-4}$ )
- $K \ll 1$      almost entirely reactants at equilibrium ( $1.00 \times 10^{-23}$ )

**Equilibrium Expression**

$$K_{\text{eq}} = \frac{[\text{C}]^c [\text{D}]^d}{[\text{A}]^a [\text{B}]^b}$$

- Law of Mass Action
- General reaction equation:  $a\text{A} + b\text{B} \leftrightarrow c\text{C} + d\text{D}$
- $K_c$  - aqueous concentrations/gases in a vessel of known volume
- $K_p$  - gases, partial pressures
- $K_w$  - water =  $1.0 \times 10^{-14}$
- $K_{\text{sp}}$  - ionic compounds (solubility product)
- $K_a$  - acids
- $K_b$  - bases
- When writing an equilibrium expression from a balanced equation, only aqueous and gaseous substances are written in the expression.
  - For example,  $\text{CaCO}_3(s) + 2\text{H}^+(aq) \leftrightarrow \text{Ca}^{2+}(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)$   
the equilibrium expression is  $K = \frac{[\text{Ca}^{2+}][\text{CO}_2]}{[\text{H}^+]^2}$

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Self-Ionization of Water

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The process of water's self-ionization is an endothermic process. Energy can be written as a reactant in the chemical equation.

Does the self-ionization increase / decrease when water is heated?



$$K_w = 1.0 \times 10^{-14} \text{ @ } 25^\circ\text{C}$$

$$K_w \ll 1$$

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$



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Equilibrium Expression

Sample WS#1 - Equilibrium Expression

1. Write the equilibrium expression for the following reaction:  $4\text{NH}_3(\text{g}) + 7\text{O}_2(\text{g}) \leftrightarrow 4\text{NO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g})$

$$K_c = \frac{[\text{NO}_2]^4 [\text{H}_2\text{O}]^6}{[\text{NH}_3]^4 [\text{O}_2]^7}$$

\* When given in M or moles and volume.

$$K_p = \frac{(P_{\text{NO}_2})^4 (P_{\text{H}_2\text{O}})^6}{(P_{\text{NH}_3})^4 (P_{\text{O}_2})^7}$$

\* When species are given in pressure units (kPa, atm, mmHg, torr, etc.)

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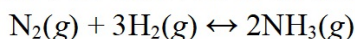
**Equilibrium Constant**

**Effect on Changing the balanced Equation of an Equilibrium Equation.**

- Multiplying the coefficients by a factor, n, raises K to the power of the factor, n,  $K^n$
- Reversing the equation - take the reciprocal of K,  $1/K$ .
- Adding equations - multiply K values.

**Affecting K Practice Problems**

2. The following equilibrium concentrations were observed for the Haber process at 127°C.



$$[\text{NH}_3] = 3.1 \times 10^{-2} \text{ mol/L}, [\text{N}_2] = 8.5 \times 10^{-1} \text{ mol/L}, \text{ and } [\text{H}_2] = 3.1 \times 10^{-3} \text{ mol/L}$$

a) Calculate the value of K at 127°C for this reaction.

$$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(3.1 \times 10^{-2})^2}{(8.5 \times 10^{-1})(3.1 \times 10^{-3})^3} = \boxed{3.8 \times 10^4}$$

$K_c \gg 1$   
\*product driven equilibrium

b) Calculate the value of the equilibrium constant at 127°C for the reaction:  $2\text{NH}_3(\text{g}) \leftrightarrow \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$

$$K_r = \frac{1}{K_f} = \frac{1}{(3.8 \times 10^4)} = \boxed{2.6 \times 10^{-5}}$$

c) Calculate the value of the equilibrium constant at 127°C for the reaction:  $1/2\text{N}_2(\text{g}) + 3/2\text{H}_2(\text{g}) \leftrightarrow \text{NH}_3(\text{g})$

$$K' = (K)^{1/2} = (3.8 \times 10^4)^{1/2} = \boxed{1.9 \times 10^2}$$

(Ans:  $3.8 \times 10^4$ ,  $2.6 \times 10^{-5}$ , and  $1.9 \times 10^2$ )



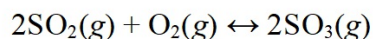
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**Equilibrium Intro**

3. The following results were collected for two experiments involving the reaction at 600°C between sulfur dioxide and oxygen to form gaseous sulfur trioxide. Show that the equilibrium constant is the same in both cases.



Experiment 1		Experiment 2	
Initial	Equilibrium	Initial	Equilibrium
$[\text{SO}_2]_0 = 2.00M$	$[\text{SO}_2] = 1.50M$	$[\text{SO}_2]_0 = 0.500M$	$[\text{SO}_2] = 0.590M$
$[\text{O}_2]_0 = 1.50M$	$[\text{O}_2] = 1.25M$	$[\text{O}_2]_0 = 0$	$[\text{O}_2] = 0.0450M$
$[\text{SO}_3]_0 = 3.00M$	$[\text{SO}_3] = 3.50M$	$[\text{SO}_3]_0 = 0.350M$	$[\text{SO}_3] = 0.260M$

Experiment # 1

$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]}$$

$$K_c = \frac{(3.5)^2}{(1.5)^2 (1.25)} = \underline{4.36}$$

Experiment #2

$$K_c = \frac{(0.26)^2}{(0.59)^2 (0.045)} = \underline{4.32}$$

4.36  $\doteq$  4.32 are very close

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**Equilibrium of Gases**

- $K_p$  uses partial pressures of gases in a sealed vessel while  $K_c$  is for concentration.
- $K_p = K_c(RT)^{\Delta n}$  where  $\Delta n$  is the change in moles of gaseous species between reactants and products.
  - $\Delta n$  = total moles of gas produced - total moles of gas reacting
  - $R$  - ideal gas law constant (0.0821)
  - $T$  in Kelvin
- $K_c = K_p$  when there is no net change in gaseous moles.

**Calculating Values of  $K_p$**

4. The reaction for the formation of nitrosyl chloride,  $2\text{NO}(g) + \text{Cl}_2(g) \leftrightarrow 2\text{NOCl}(g)$ , was studied at  $25^\circ\text{C}$ . The pressures at equilibrium were found to be  $P_{\text{NOCl}} = 1.2\text{atm}$ ,  $P_{\text{NO}} = 5.0 \times 10^{-2}\text{atm}$ , and  $P_{\text{Cl}_2} = 3.0 \times 10^{-1}\text{atm}$ . Calculate the value of  $K_p$  for this reaction at  $25^\circ\text{C}$ .  $\Delta n = 2 - (2+1) = -1$  (Ans:  $1.9 \times 10^3$ )

GvN

$$P_{\text{NOCl}} = 1.2\text{atm}$$
$$P_{\text{NO}} = 5.0 \times 10^{-2}\text{atm}$$
$$P_{\text{Cl}_2} = 3.0 \times 10^{-1}\text{atm}$$

$$K_c = \frac{N_T K}{(P_{\text{NO}})^2 (P_{\text{Cl}_2})}$$

$$= \boxed{1.9 \times 10^3}$$

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**Equilibrium of Gases**

- $K_p$  uses partial pressures of gases in a sealed vessel while  $K_c$  is for concentration.
  - $K_p = K_c(RT)^{\Delta n}$  where  $\Delta n$  is the change in moles of gaseous species between reactants and products.
    - $\Delta n$  = total moles of gas produced - total moles of gas reacting
    - $R$  - ideal gas law constant (0.0821)
    - $T$  in Kelvin
  - $K_c = K_p$  when there is no net change in gaseous moles.
5. Using the value of  $K_p$  obtained in Sample 4, calculate the value of  $K_c$  at 25°C for the reaction
- $$2\text{NO}(g) + \text{Cl}_2(g) \leftrightarrow 2\text{NOCl}(g)$$

$$\Delta n = 2 - (2+1) = -1$$

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

$$K_c = \frac{1.9 \times 10^3}{[(0.0821)(298)]^{-1}} = \boxed{4.6 \times 10^4}$$

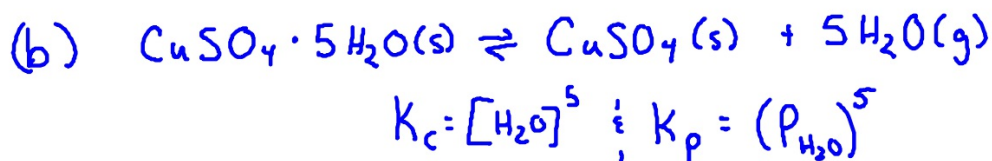
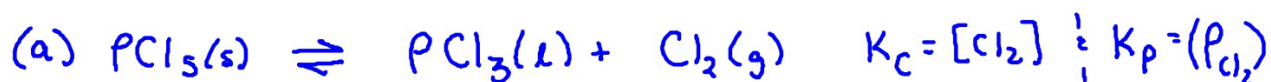
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**Equilibrium of Gases**

6. Write the expressions  $K$  and  $K_p$  for the following processes:
- The decomposition of solid phosphorus pentachloride to liquid phosphorus trichloride and chlorine gas.
  - Deep blue solid copper (II) sulfate pentahydrate is heated to drive off water vapor to form white solid copper (II) sulfate.



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**Equilibrium Review**

**Can you...**

1. ...write an equilibrium constant expression?
2. ...tell how  $K$  is changed if the stoichiometric coefficients are changed on an equation?
3. ...tell how to find  $K$  for a summary equation?
4. ...tell how  $K$  depends on the way equilibrium concentrations are expressed and how to convert  $K$  in terms of  $K_c$  vs.  $K_p$ ?
5. ...explain what  $K$  is telling you about a reaction?

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**Reaction Quotient**

**The Reaction Quotient ( $Q$ )**

- Used to determine whether a system is at equilibrium or not.
- Compare  $Q$  to  $K$  to determine if the reaction is at equilibrium or will shift left or right.
- For the reaction,  $aA + bB \leftrightarrow cC + dD$

$$Q_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

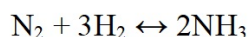
- $Q = K$  then system is at equilibrium
- $Q < K$  then system will shift right (increase amount of products (numerator))
- $Q > K$  the system will shift left (increase amount of reactants (denominator))

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Reaction Quotient

Using Q (Reaction Quotient) to Determine if Reaction is at Equilibrium

7. For the synthesis of ammonia at 500°C, the equilibrium constant is  $6.0 \times 10^{-2}$ . Predict the direction in which the system will shift to reach equilibrium in each of the following cases.



$$K = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = 6.0 \times 10^{-2}$$

- a.  $[\text{NH}_3]_0 = 1.0 \times 10^{-3} M$ ,  $[\text{N}_2]_0 = 1.0 \times 10^{-5} M$ , and  $[\text{H}_2]_0 = 2.0 \times 10^{-3} M$   
b.  $[\text{NH}_3]_0 = 2.00 \times 10^{-4} M$ ,  $[\text{N}_2]_0 = 1.50 \times 10^{-5} M$ , and  $[\text{H}_2]_0 = 3.54 \times 10^{-1} M$   
c.  $[\text{NH}_3]_0 = 1.0 \times 10^{-4} M$ ,  $[\text{N}_2]_0 = 5.0 M$ , and  $[\text{H}_2]_0 = 1.0 \times 10^{-2} M$  (Ans: shifts left, no shift, shifts right)

(a)  $Q = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(1 \times 10^{-3})^2}{(1.0 \times 10^{-5})(2.0 \times 10^{-3})^3} = 1.3 \times 10^7$      $6.0 \times 10^{-2}$

$Q > K$   
← shift left

(b)  $Q = \frac{(2.00 \times 10^{-4})^2}{(1.5 \times 10^{-5})(3.54 \times 10^{-1})^3} = 6.0 \times 10^{-2}$

$Q = K$  (no shift)

(c)  $Q = \frac{(1.0 \times 10^{-4})^2}{(5)(1 \times 10^{-2})^3} = 2.0 \times 10^{-3}$      $6.0 \times 10^{-2}$

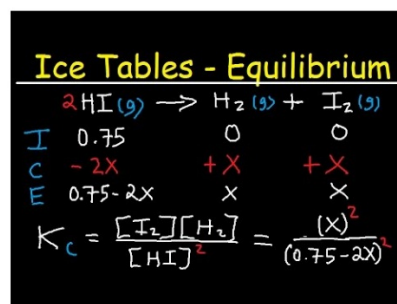
$Q < K$   
(shifts right)

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**Equilibrium RICE Tables**

**General Steps in Solving Equilibrium Problems**

- Set up a RICE table
  - R - balanced reaction of reacting species
  - I - initial concentration of species
  - C - determine the change occurring in terms of  $x$
  - E - express the equilibrium concentrations in terms of  $x$
- Set up equilibrium expression and set it equal to a value (if given).
- If given a  $K$  value, use it to solve for  $x$ .
- If  $K$  value is VERY Small (less than  $1.0 \times 10^{-4}$ ) then you can eliminate the change to the initial concentration.
  - Hints: (a) Look for very small  $K$  values, " $x$ " may be negligible.
  - (b) If " $x$ " is large enough, then you must subtract it from the initial concentration. Your math will be simplified if the problem is a perfect square. If not, you must use the quadratic formula.
  - (c) If none of the initial concentrations are zero, then  $Q$  must be calculated first to determine direction of the shift before RICE table.



Reaction	$2\text{HI}(g) \leftrightarrow \text{H}_2(g) + \text{I}_2(g)$		
Initial	0.50	0	0
Change	-2x	+x	+x
Equilibrium	0.50-2x	0+x	0+x

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{(x)(x)}{(0.5-2x)^2} = \frac{(x)^2}{(0.5-2x)^2}$$

\*If  $K \ll 1$ , then the -2x can be relegated to zero. What this means is the change is sooo small that it does not affect the initial [HI].



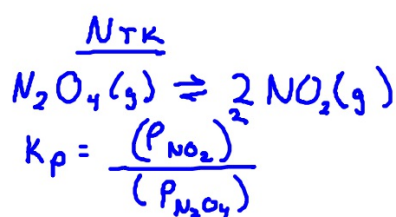
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**Equilibrium RICE Tables**

**Calculating Equilibrium Pressures Practice Problems**

8. Dinitrogen tetroxide, in its liquid state, was used as one of the fuels on the lunar lander for the NASA Apollo missions. In its gas phase it decomposes to gaseous nitrogen:  $\text{N}_2\text{O}_4(\text{g}) \leftrightarrow 2\text{NO}_2(\text{g})$ . Consider an experiment in which gaseous  $\text{N}_2\text{O}_4$  was placed in a flask and allowed to reach equilibrium at a temperature where  $K_p = 0.133$ . At equilibrium, the pressure of  $\text{N}_2\text{O}_4$  was found to be 2.71 atm. Calculate the equilibrium pressure of  $\text{NO}_2(\text{g})$ . (Ans: 0.600 atm)

Giv  
 $K_p = 0.133$   
 $P_{\text{N}_2\text{O}_4} = 2.71 \text{ atm}$



$P_{\text{NO}_2} = \underline{0.600 \text{ atm}}$

SOLVE:

$$K_p = 0.133 = \frac{(P_{\text{NO}_2})^2}{(P_{\text{N}_2\text{O}_4})}$$

$$(0.133)(2.71) = (P_{\text{NO}_2})^2$$

$$P_{\text{NO}_2} = \sqrt{(0.133)(2.71)} = \boxed{0.600 \text{ atm}}$$

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**Equilibrium RICE Tables**

9. At a certain temperature a 1.00L flask initially contained 0.298mol  $\text{PCl}_3(\text{g})$  and  $8.70 \times 10^{-3}$  mol  $\text{PCl}_5(\text{g})$ . After the system had reached equilibrium,  $2.00 \times 10^{-3}$  mol  $\text{Cl}_2(\text{g})$  was found in the flask. Gaseous  $\text{PCl}_5$  decomposes according to the reaction:  $\text{PCl}_5(\text{g}) \leftrightarrow \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$ . Calculate the equilibrium concentrations of all species and the value of  $K$ .

(Ans:  $[\text{Cl}_2] = 2.00 \times 10^{-3} \text{ M}$ ,  $[\text{PCl}_3] = 0.300 \text{ M}$ ,  $[\text{PCl}_5] = 6.70 \times 10^{-3} \text{ M}$ , and  $K = 8.96 \times 10^{-2}$ )

Given  
 $[\text{PCl}_3]_0 = \frac{0.298 \text{ mol}}{1 \text{ L}} = 0.298 \text{ M}$

$$[\text{PCl}_5]_0 = \frac{8.70 \times 10^{-3} \text{ mol}}{1 \text{ L}} = 8.70 \times 10^{-3} \text{ M}$$

$$[\text{Cl}_2]_E = \frac{2.00 \times 10^{-3}}{1 \text{ L}} = 2.00 \times 10^{-3} \text{ M} = x$$

$$[\text{PCl}_3]_E = 0.298 + x = 0.298 + 0.002 = 0.300 \text{ M}$$

$$[\text{PCl}_5]_E = 8.70 \times 10^{-3} - x = 0.0087 - 0.002 = 0.0067 = 6.7 \times 10^{-3} \text{ M}$$

R	$\text{PCl}_5(\text{g})$	$\rightleftharpoons$	$\text{PCl}_3(\text{g})$	$+$	$\text{Cl}_2(\text{g})$
I	$8.70 \times 10^{-3}$		0.298		0
C	$-x$		$+x$		$+x$
E	$8.70 \times 10^{-3} - x$		$0.298 + x$		$x$

$$K_c = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]} = \frac{(0.3)(0.002)}{(0.0067)} = 8.96 \times 10^{-2}$$

Equilibrium RICE Tables

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10. Carbon monoxide reacts with steam to produce carbon dioxide and hydrogen. At 700K, the equilibrium constant is 5.10. Calculate the equilibrium concentrations of all species if 1.000mol of each component is mixed in a 1.000L flask.

→ (Ans:  $[CO] = [H_2O] = 0.613M$  and  $[CO_2] = [H_2] = 1.387M$ )

$K_c = 5.10$	<u>Gun</u>	R	CO	+ H <sub>2</sub> O	⇌	CO <sub>2</sub>	+ H <sub>2</sub>	
		I	1.0	1.0		1.0	1.0	
		C	-x	-x		+x	+x	
		E	1-x	1-x		1+x	1+x	

$[CO] = 1 - 0.387 = 0.613 M$   
 $[H_2O] = 1 - 0.387 = 0.613 M$   
 $[CO_2] = 1 + 0.387 = 1.387 M$   
 $[H_2] = 1 + 0.387 = 1.387 M$

check Q

$$Q = \frac{[CO][H_2O]}{[CO_2][H_2]} = \frac{(1)(1)}{(1)(1)} = 1$$

Q vs K  
 $1 < 5.10$   
(shifts right)

$$K = \frac{[CO_2][H_2]}{[CO][H_2O]}$$

Solve:

$$K = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{(1+x)(1+x)}{(1-x)(1-x)} = \sqrt{\frac{(1+x)^2}{(1-x)^2}} = \sqrt{5.10}$$

$$\frac{(1+x)}{(1-x)} = 2.2583$$

$$1+x = 2.2583 - 2.2583x$$

$$3.2583x = 2.2583 - 1$$

$$x = \frac{1.2583}{3.2583} = \boxed{0.386}$$

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**Equilibrium RICE Tables**

11. Assume that the reaction for the formation of gaseous hydrogen fluoride from hydrogen and fluorine has an equilibrium constant of  $1.15 \times 10^2$  at a certain temperature. In a particular experiment, 3.000 mol of each component was added to a 1.500 L flask. Calculate the equilibrium concentrations of all species.

(Ans:  $[H_2] = [F_2] = 0.472 M$  and  $[HF] = 5.056 M$ )

GvN  
 $K_c = 1.15 \times 10^2$   
 $[HF] = \frac{3.0}{1.5} = 2.000 M$   
 $[H_2] = \frac{3}{1.5} = 2.000 M$   
 $[F_2] = \frac{3}{1.5} = 2.000 M$   
 Check(Q)  
 $Q = \frac{(HF)^2}{(H_2)(F_2)}$   
 $= \frac{(2)^2}{(2)(2)} = 1$   
 (shifts right) Q vs K  
 $1 < 1.15 \times 10^2$

NTK

R	$H_2 + F_2 \rightleftharpoons 2HF$
I	2    2    2
C	-x   -x   +2x
E	2-x   2-x   2+2x

UNK

$[H_2] =$	$\frac{0.471}{1.5}$	M
$[F_2] =$	$\frac{0.471}{1.5}$	M
$[HF] =$	$\frac{5.056}{1.5}$	M

Solve:  $K = \frac{[HF]^2}{[H_2][F_2]} = \frac{(2+2x)^2}{(2-x)(2-x)}$   
 (x):  $\sqrt{1.15 \times 10^2} = \sqrt{\frac{(2+2x)^2}{(2-x)^2}}$   
 $10.723 = \frac{2+2x}{2-x} \rightarrow 21.448 \cdot 2 = 2x + 10.723x$   
 $10.723(2-x) = 2+2x$   
 $21.448 - 10.723x = 2+2x$   
 $19.448 = 12.723x$   
 $x = \frac{19.448}{12.723}$   
 $x = 1.529$

$[H_2] = [F_2] = 2 - x$   
 $= 2 - 1.529 = 0.471$   
 $[H_2] = [F_2] = 0.471$   
 $[HF] = 2 + 2(x)$   
 $[HF] = 2 + 2(1.529)$   
 $[HF] = 5.058 M$

---

**Equilibrium**

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**Topic#8****Equilibrium RICE Tables**

12. (OYO) Assume gaseous hydrogen iodide is synthesized from hydrogen gas and iodine vapor at a temperature where the equilibrium constant  $1.00 \times 10^2$ . Suppose HI at  $5.000 \times 10^{-1}$  atm,  $H_2$  at  $1.000 \times 10^{-2}$  atm, and  $I_2$  at  $5.000 \times 10^{-3}$  atm are mixed in a 5.000L flask. Calculate the equilibrium pressure of all species.

(Ans:  $P_{HI} = 4.29 \times 10^{-1}$  atm,  $P_{H_2} = 4.55 \times 10^{-2}$  atm, and  $P_{I_2} = 4.05 \times 10^{-2}$  atm)

---

**Equilibrium**  
**Topic#8**

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**LeChatelier's Principle**

**Le Chatelier's Principle**

- A system at equilibrium when stressed will compensate for the stress and return to equilibrium.
- Stresses include addition/removal of heat (change in  $T$ ), change in pressure (changed by changing volume), change in concentration of a reactant or product.
  - an addition of a species/heat will force the reaction to shift away from the stress in an effort to re-establish equilibrium.
  - a removal of a species/heat will force the reaction to shift towards the stress in an effort to re-establish equilibrium.
  - a decrease in volume will force the reaction to shift towards the side with the least total gaseous moles in an effort to re-establish equilibrium.
  - an increase in volume will force the reaction to shift towards the side with the most total gaseous moles in an effort to re-establish equilibrium.
- the value of  $K_{eq}$  can ONLY be changed by a change in temperature. Each value of  $K$  is dependent on the temperature of the reaction.

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**LeChatelier's Principle**

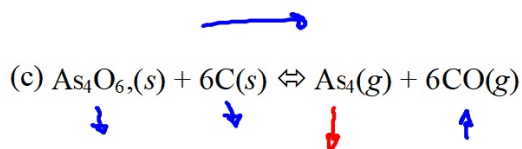
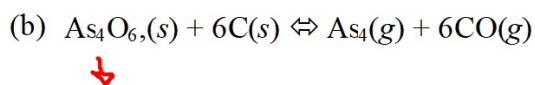
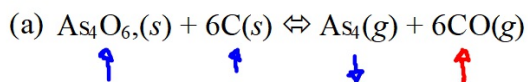
**Equilibrium**  
**Topic#8**  
**(Topic)**

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13. Arsenic can be extracted from its ores by first reacting the ore with oxygen (called *roasting*) to form solid  $\text{As}_4\text{O}_6$ , which is then reduced using carbon.  $\text{As}_4\text{O}_6(s) + 6\text{C}(s) \rightleftharpoons \text{As}_4(g) + 6\text{CO}(g)$   
Predict the direction of the shift of the equilibrium position in response to each of the following changes in conditions.

- (a) Addition of carbon monoxide
- (b) Addition or removal of carbon or tetraarsenic hexoxide ( $\text{As}_4\text{O}_6$ )
- (c) Removal of gaseous arsenic ( $\text{As}_4$ )

$$K = [\text{As}_4][\text{CO}]^6$$



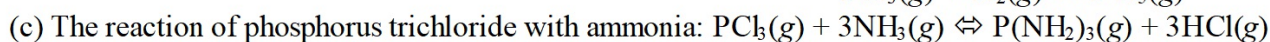
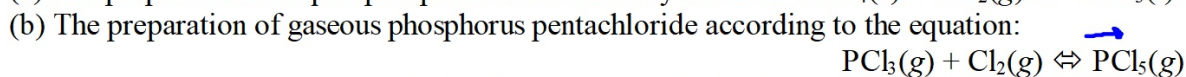
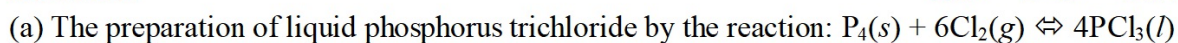
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LeChatelier's Principle

Equilibrium  
Topic#8  
(Topic)

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14. Predict the shift in equilibrium position that will occur for each of the following processes when the volume is reduced.

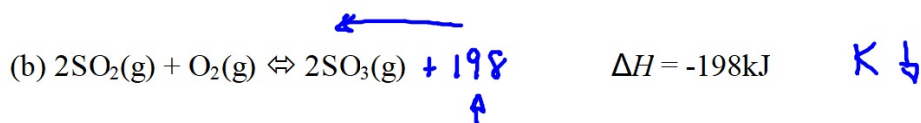
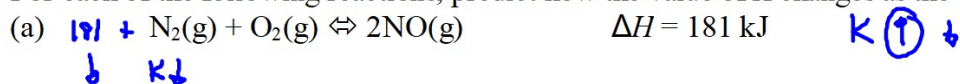


6 mol → 0 mol

→

no shift

15. For each of the following reactions, predict how the value of  $K$  changes as the temperature is increased.



b K ↓

←

A



**Equilibrium**

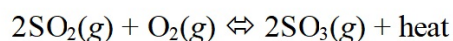
**Topic#8**

**LeChatelier's Principle**

key: --- unchanged    ↑ increases    ↓ decreases

16. (OYO) For the reaction,  $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g}) + \text{heat}$ , identify the direction the reaction will shift to alleviate the stress. Determine whether each of the species in the reaction will increase/decrease in concentration.

$$K_{\text{eq}} = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}$$



	shift	[SO <sub>2</sub> ]	[O <sub>2</sub> ]	[SO <sub>3</sub> ]	K
a. Addition of O <sub>2</sub> .					
b. Addition of SO <sub>3</sub> .					
c. Decrease in temperature.					
d. Increase in temperature.					
e. Decrease in volume of container.					
f. Increase in the volume of the container.					
g. Addition of SO <sub>2</sub> .					
h. Removal of SO <sub>3</sub> .					
i. Removal of O <sub>2</sub> .					
j. Addition of CO <sub>2</sub>					

→    ↓ ↑ ↓    ← ---    ↓ ↓ ↑    ↓ ↑ ---    ---

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Solubility  
Topic#8

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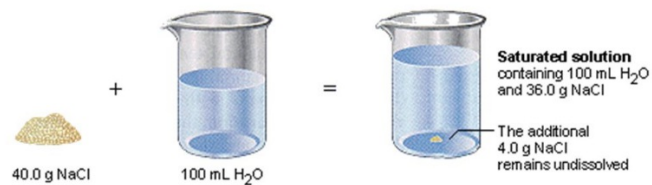
# **Solubility Equilibrium**

## **$K_{sp}$**

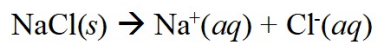
Solubility  
Topic#8

Equilibrium Intro

Saturated vs. Unsaturated Solutions

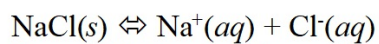


Unsaturated Solution



\*below saturation point, all of the solute dissolves, no equilibrium

Saturated Solution



\*an equilibrium has been established between the ions and solid

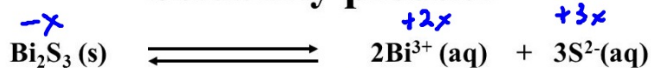
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**Formulas**

**Solubility  
Topic#8**

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**Solubility product**



$$K_{sp} = [\text{Bi}^{3+}]^2 [\text{S}^{2-}]^3$$

Solubility product constant

Molar solubility of the ions

Remember that we are dealing with molar solubilities and not concentration.

For a **saturated solution**, molar solubility is equal to molar concentration.

\*For every 1M of  $\text{Bi}_2\text{S}_3(s)$  that dissolves 2M of  $\text{Bi}^{3+}(aq)$  and 3M of  $\text{S}^{2-}(aq)$  are created.

$$K_{sp} = (2x)^2 (3x)^3$$
$$(4x^2)(27x^3)$$
$$\textcircled{108x^5}$$

$$K = (x)(x) = x^2$$

$$K = (2x)^2(x) = 4x^3$$

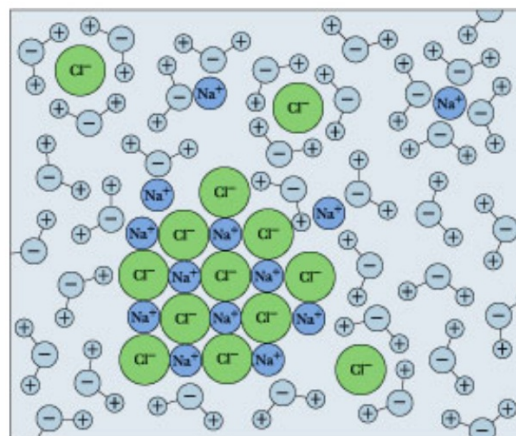
$$K = (3x)^3(x) = 27x^4$$

$$K = (3x)^3(2x)^2 = 108x^5$$

## Solubility Topic#8

### Types of Solutions

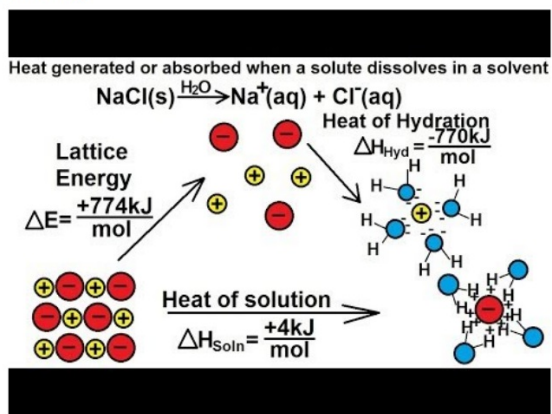
- Even though we determine a compound to be insoluble, it does have a solubility, but it is very low.
- Solubility is defined as "greater than 3 grams per 100mL of water."
- A solution must be saturated to do solubility equilibrium problems.
  - A **saturated** solution must have some of the solute at the bottom of the container.
  - An **unsaturated** solution has no solute on the bottom of the container.
  - A **supersaturated** solution has more solute per 100mL of water than it should at the specified temperature. This type of solution also has no solute at the bottom. Process: heat solvent (water) and create a saturated solution at this high temperature. Pour of supernatant (water above solid) and then cool slowly. The new cooled solution is supersaturated.



### Animation of Dissolution

<http://www.dynamicscience.com.au/tester/solutions1/chemistry/chemicaldemos/molvie1.swf>

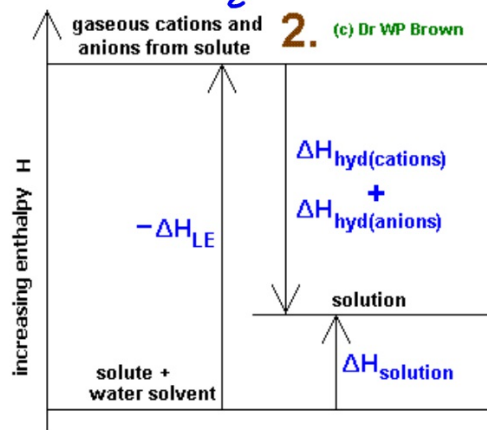
## Solubility



Compound	Lattice Energy (kJ/mol)	Hydration Energy (kJ/mol)	Solubility in Water g/100 mL
LiCl	834	-884	63.7 (0 °C)
NaCl	774	-760	35.7 (0 °C)
KCl	701	-685	34.7 (20 °C)
RbCl	680	-664	77 (0 °C)
CsCl	657	-640	162 (0 °C)
LiF	1032	-1005	0.3 (18 °C)
KF	813	-819	92.3 (18 °C)
KI	632	-617	127 (0 °C)
RbF	776	-792	130.6 (18 °C)
SrCl <sub>2</sub>	2110	-2161	53.8
AgCl	916	-851	8.9 x 10 <sup>-5</sup>

## Solubility Topic#8

$$F_c = \frac{k \delta_1 \delta_2}{d}$$

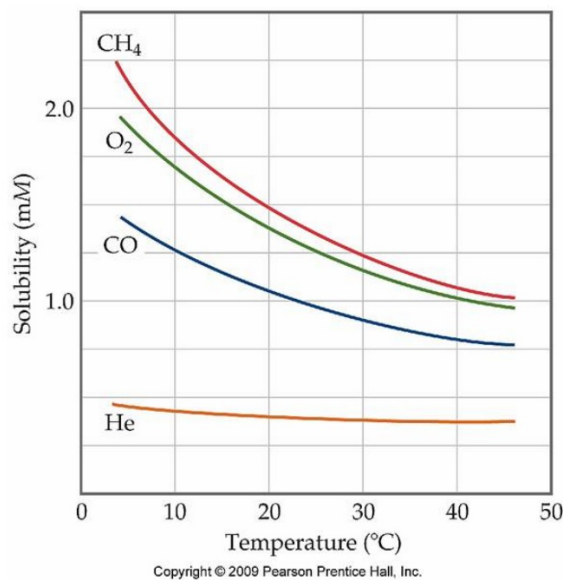
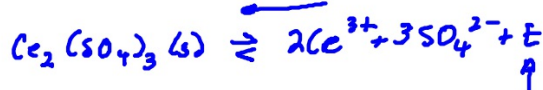
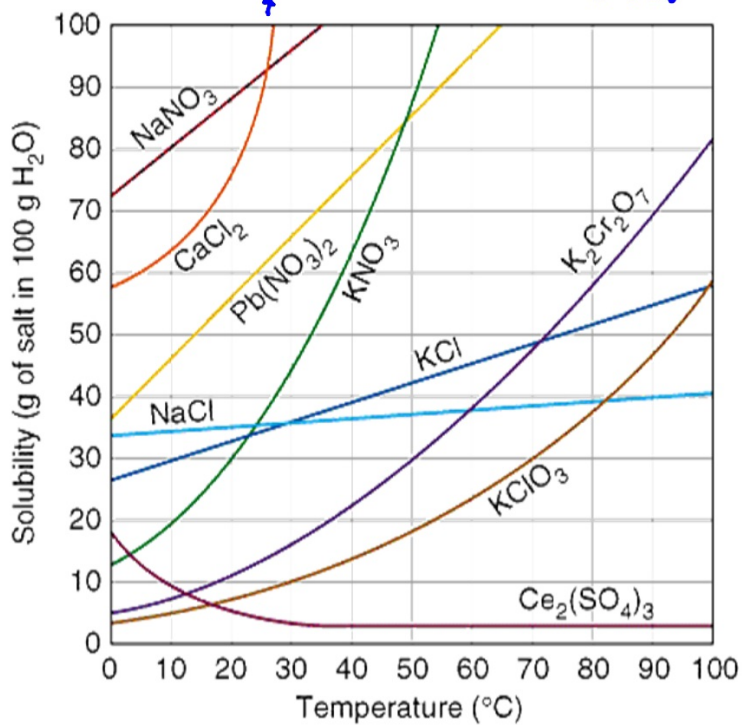
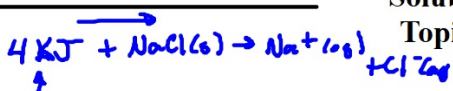


$$\Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

$$= 774 + -770 = +4 \text{ kJ/mol}$$

**Solubility**

**Solubility Topic#8**

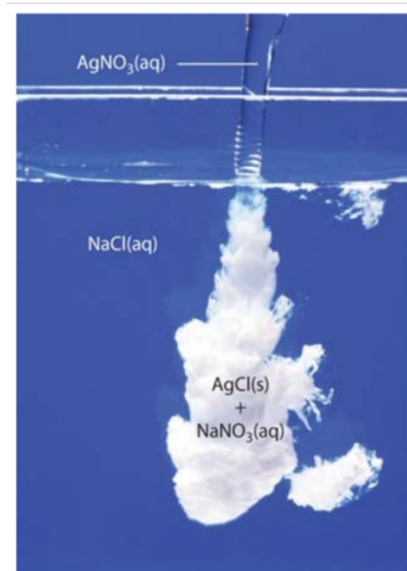


**TABLE 11.2** Some Enthalpies and Entropies of Solution in Water at 25°C

## Solubility Topic#8

### Solubility

- Slightly soluble salts create a dynamic equilibrium in a saturated solution.
  - For Example: a solution of  $\text{AgNO}_3$  and  $\text{NaCl}$  are mixed creating  $\text{AgCl}$  (generally, all silver (I) compounds are insoluble).
    - solid  $\text{AgCl}$  is formed, no ions initially present
    - as the salt dissolves, cations ( $\text{Ag}^+$ ) and anions ( $\text{Cl}^-$ ) are created
    - $[\text{Ag}^+]$  and  $[\text{Cl}^-]$  increase until equilibrium is established (a saturated solution)
    - $K_{\text{sp}}$  is the product of the ions in solution.
    - For a saturated solution of  $\text{AgCl}$  the net ionic equation is
$$\text{AgCl}(s) \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)$$
    - $K_{\text{sp}} = [\text{cation}]^n[\text{anion}]^m$  where  $n$  and  $m$  are the coefficients of the ions in solution (net ionic equation)
$$K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10} \text{ (very low indeed!)}$$
    - $\text{AgCl}(s)$  is not appear in solubility equilibrium expression.
- I have included a table of solubilities in the reference section of the Equilibrium/Solubility Equilibrium page on our website.



$$x^2 = 1.8 \times 10^{-10}$$
$$x = 1.3 \times 10^{-5} M$$



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**Solubility**  
**Topic#8**

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**Solubility**

**Solubility Rules**

- (1) nitrate salts are ALL soluble
- (2) alkali metal and ammonium salts are ALL soluble
- (3) chlorides, bromides, and iodides salts are mostly soluble (Dirty 3:  $\text{Ag}^+$ ,  $\text{Pb}^{2+}$ , and  $\text{Hg}_2^{2+}$ )
- (4) sulfate salts are mostly soluble (Except  $\text{BaSO}_4$ ,  $\text{CaSO}_4$ ,  $\text{PbSO}_4$ , and  $\text{Hg}_2\text{SO}_4$ )
- (5) hydroxides are mostly insoluble (Except:  $\text{NH}_4\text{OH}$ ,  $\text{LiOH}$ ,  $\text{NaOH}$ ,  $\text{KOH}$ ,  $\text{RbOH}$ ,  $\text{CsOH}$ ,  $\text{Ca}(\text{OH})_2$ ,  $\text{Sr}(\text{OH})_2$ , and  $\text{Ba}(\text{OH})_2$ )
- (6) sulfides ( $\text{S}^{2-}$ ), carbonates ( $\text{CO}_3^{2-}$ ), chromates ( $\text{CrO}_4^{2-}$ ), and phosphates ( $\text{PO}_4^{3-}$ ) are insoluble

**Equilibrium Sample WS#2 - Solubility Equilibrium**

**Solubility Practice Problems**

1. Determine the solubility of each of the following compounds when placed in water. Write a yes or no and the number of the rule you used to come to your conclusion.

a.  $\text{KBr}$

d. sodium acetate

b.  $\text{PbCO}_3$

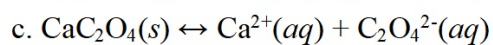
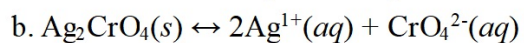
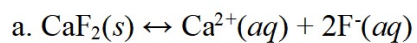
e. silver (I) iodide

c. zinc (II) hydroxide

f. cadmium (II) sulfide

**Solubility Expression ( $K_{sp}$ ) Practice Problems**

2. Write the  $K_{sp}$  expression for each of the following solubility equations:



(a.)  $K_{sp} = [\text{Ca}^{2+}][\text{F}^{-}]^2$

(b.)  $K_{sp} = [\text{Ag}^{+}]^2[\text{CrO}_4^{2-}]$

(c.)  $K_{sp} = [\text{Ca}^{2+}][\text{C}_2\text{O}_4^{2-}]$

Solubility  
Topic#8

Solubility Expression ( $K_{sp}$ )

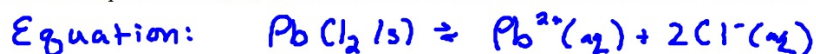
**Determining  $K_{sp}$  from Experimental Measurements.**

- Use RICE table to determine the amount of each ion after equilibrium has been established.
- Substitute concentrations and solve for  $K_{sp}$ .

$K_{sp}$  Value Practice Problems

3. Calculate the  $K_{sp}$  of lead (II) chloride if the lead concentration is  $1.62 \times 10^{-2} M$ .

(Ans:  $1.70 \times 10^{-5}$ )



Change:  $-x$                        $+x$                        $+2x$

$$[Pb^{2+}] = x = \frac{1.62 \times 10^{-2} M}{1}$$

$$[Cl^{-}] = 2x = 2(1.62 \times 10^{-2}) \\ = 0.0324 M$$

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

$$K_{sp} = (1.62 \times 10^{-2})(0.0324)^2 \\ = 1.7 \times 10^{-5}$$

---

**Solubility Expression ( $K_{sp}$ )**

**Solubility  
Topic#8**

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**Determining  $K_{sp}$  from Experimental Measurements.**

- Use RICE table to determine the amount of each ion after equilibrium has been established.
- Substitute concentrations and solve for  $K_{sp}$ .

4. (OYO) Calculate the  $K_{sp}$  value for bismuth (III) sulfide, which has a solubility of  $1.0 \times 10^{-15}$  mol/L at  $25^\circ\text{C}$ .  
(Ans:  $1.1 \times 10^{-73}$ )

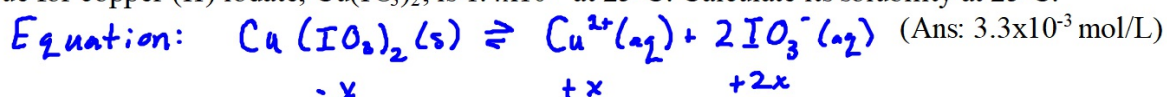
Solubility  
Topic#8

Solubility Expression ( $K_{sp}$ )

Estimating Salt Solubility from  $K_{sp}$ .

- Set up  $x$  as the amount of salt that dissolves.
- Use  $-x$  and  $+x$  in RICE table to solve the amounts of the cations and anions in solution.

5. The  $K_{sp}$  value for copper (II) iodate,  $\text{Cu}(\text{IO}_3)_2$ , is  $1.4 \times 10^{-7}$  at  $25^\circ\text{C}$ . Calculate its solubility at  $25^\circ\text{C}$ .



$$K_{sp} = [\text{Cu}^{2+}][\text{IO}_3^-]^2 = 1.4 \times 10^{-7}$$

$$(x)(2x)^2 = 1.4 \times 10^{-7}$$

$$4x^3 = 1.4 \times 10^{-7}$$

$$x = \sqrt[3]{\frac{1.4 \times 10^{-7}}{4}} = \boxed{3.3 \times 10^{-3} \text{ M}}$$

$x$  is molar solubility

$$\frac{3.3 \times 10^{-3} \text{ mol}}{\text{L}} \times \frac{413.35 \text{ g}}{\text{mol}} = 3.3 \times 10^{-3} \frac{\text{mol}}{\text{L}} \times \frac{413.35 \text{ g}}{\text{mol}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \div 10$$

---

**Solubility**  
**Topic#8**

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**Solubility Expression ( $K_{sp}$ )**

**Estimating Salt Solubility from  $K_{sp}$ .**

- Set up  $x$  as the amount of salt that dissolves.
- Use  $-x$  and  $+x$  in RICE table to solve the amounts of the cations and anions in solution.

6. (OYO) The  $K_{sp}$  value for cobalt (II) phosphate,  $\text{Co}_3(\text{PO}_4)_2$ , is  $1.4 \times 10^{-37}$  at  $25^\circ\text{C}$ . Calculate its solubility at this temperature. At equilibrium, what are the concentrations of each ion?

(Ans:  $1.7 \times 10^{-8} \text{M}$ ,  $[\text{Co}^{2+}] = 5.0 \times 10^{-8} \text{M}$ , and  $[\text{PO}_4^{3-}] = 3.3 \times 10^{-8} \text{M}$ )

Solubility  
Topic#8

Solubility Expression ( $K_{sp}$ )

**Estimating Salt Solubility from  $K_{sp}$ .**

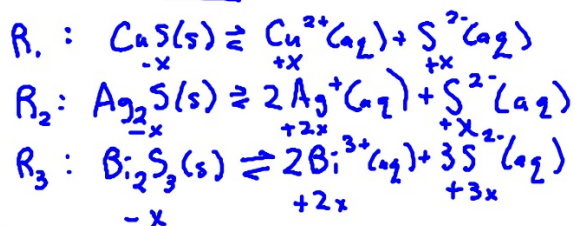
- Set up  $x$  as the amount of salt that dissolves.
- Use  $-x$  and  $+x$  in RICE table to solve the amounts of the cations and anions in solution.

7. Order the following salts based on their relative solubilities: CuS ( $K_{sp} = 8.5 \times 10^{-45}$ ),  $Ag_2S$  ( $K_{sp} = 1.6 \times 10^{-49}$ ), and  $Bi_2S_3$  ( $K_{sp} = 8.1 \times 10^{-73}$ ) (Ans: (most)  $Bi_2S_3 > Ag_2S > CuS$  (least))

GUN

$K_{sp}$ 's of CuS,  $Ag_2S$ , &  
 $Bi_2S_3$

NTK



UNK

order:  $Bi_2S_3 > Ag_2S > CuS$

$$1.5 \times 10^{-15} > 3.4 \times 10^{-17} > 9.2 \times 10^{-23}$$

$Bi_2S_3 > Ag_2S > CuS$

SOLVE:  
Solve for  $x$

$$R_1: K_{sp} = (x)(x) = x^2$$

$$(CuS) \quad x^2 = 8.5 \times 10^{-45}$$

$$x = \sqrt{8.5 \times 10^{-45}}$$

$$x = \boxed{9.2 \times 10^{-23} M}$$

$$R_2: K_{sp} = (2x)^2(x)$$

$$(Ag_2S) \quad 4x^3 = 1.6 \times 10^{-49}$$

$$x = \sqrt[3]{\frac{1.6 \times 10^{-49}}{4}}$$

$$x = \boxed{3.4 \times 10^{-17} M}$$

$$R_3: K_{sp} = (2x)^2(3x)^3$$

$$(Bi_2S_3) \quad 108x^5 = 8.1 \times 10^{-73}$$

$$x = \sqrt[5]{\frac{8.1 \times 10^{-73}}{108}}$$

$$x = 1.5 \times 10^{-15}$$

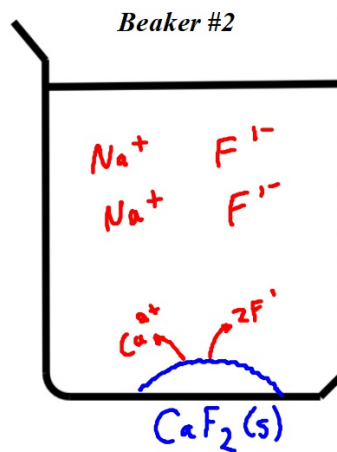
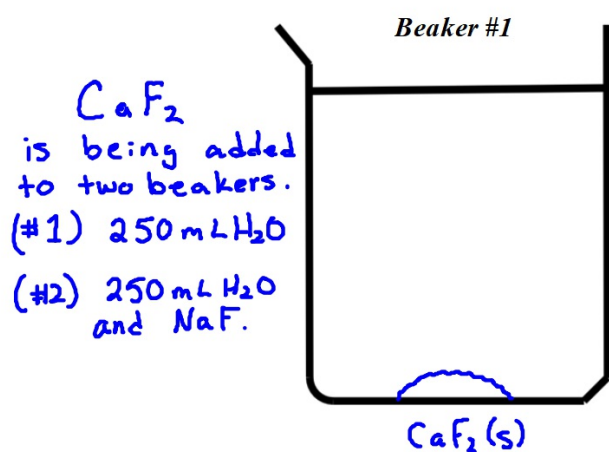
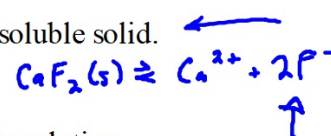
Solubility  
Topic#8

Solubility Expression ( $K_{sp}$ ) and Common Ion

$$K_{sp} = [Ca^{2+}][F^-]^2$$

Common Ion using  $K_{sp}$ .

- The solvent already has ions in it. One of which is a common ion with the insoluble solid.
- Set up  $x$  as the amount of salt that dissolves.
- The initial amount for at least one of the ions has a value other than zero
- Use  $-x$  and  $+x$  in RICE table to solve the amounts of the cations and anions in solution.



Since beaker (2) has  $F^-$  ions already in solution, the amount of  $CaF_2$  that will dissolve decreases. So,  $[Ca^{2+}]$  in beaker (2) will be lower than  $[Ca^{2+}]$  in beaker (1). This is called the common ion effect.



Solubility  
Topic#8

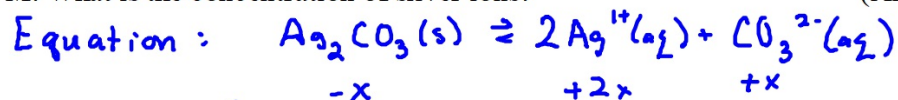
Solubility Expression ( $K_{sp}$ ) and Common Ion

**Common Ion using  $K_{sp}$ .**

- The solvent already has ions in it. One of which is a common ion with the insoluble solid.
- Set up  $x$  as the amount of salt that dissolves.
- The initial amount for at least one of the ions has a value other than zero
- Use  $-x$  and  $+x$  in RICE table to solve the amounts of the cations and anions in solution.

**Common Ion Practice Problems**

8. The  $K_{sp}$  for silver carbonate is  $8.4 \times 10^{-12}$ . The concentration of carbonate ions in a saturated solution is  $1.28 \times 10^{-4} M$ . What is the concentration of silver ions? (Ans:  $2.56 \times 10^{-4} M$ )



$$[\text{CO}_3^{2-}] = x = 1.28 \times 10^{-4} M$$

$$[\text{Ag}^+] = 2x = 2(1.28 \times 10^{-4}) = 2.56 \times 10^{-4} M$$



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**Solubility Expression ( $K_{sp}$ ) and Common Ion**      **Solubility**  
**Topic#8**

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10. (OYO) The  $K_{sp}$  for  $\text{Ca}(\text{OH})_2$  is known to be  $4.68 \times 10^{-6}$ . What is the solubility of  $\text{Ca}(\text{OH})_2$  in  $0.0860 \text{ M}$   $\text{Ba}(\text{OH})_2$ ? (Ans:  $1.58 \times 10^{-4} \text{ M}$ )

Solubility  
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Precipitation Reactions

With some knowledge of the reaction quotient ( $Q$ ), we can decide

- $Q < K_{sp}$ , the system is not at equilibrium (*unsaturated*)
- $Q = K_{sp}$ , the system is at equilibrium (*saturated*)
- $Q > K_{sp}$ , the system is not at equilibrium (*supersaturated*)

*Precipitates form when the solution is supersaturated!!!!*

$K_{sp}$  vs.  $Q$  Practice Problems

11. A solution is prepared by adding 750.0mL of  $4.00 \times 10^{-3} M$   $Ce(NO_3)_3$  to 300.0mL of  $2.00 \times 10^{-2} M$   $KIO_3$ . Will  $Ce(IO_3)_3$  ( $K_{sp} = 1.9 \times 10^{-10}$ ) precipitate from this solution? (Ans: yes;  $Q > K_{sp}$ )

$Ce^{3+}$   $IO_3^-$

GUN

750 mL  $\times$   $4.00 \times 10^{-3} M$   $Ce^{3+}$   $\rightarrow$   $3.0 \text{ mmol } Ce^{3+}$

300 mL  $\times$   $2.00 \times 10^{-2} M$   $IO_3^-$   $\rightarrow$   $6 \text{ mmol } IO_3^-$

NTK

$$Ce(IO_3)_3(s) \rightleftharpoons Ce^{3+}(aq) + 3 IO_3^-(aq)$$

-x +x +3x

SOLVE:  $Q = (Ce^{3+})(IO_3^-)^3$

$Q = \left(\frac{3.0 \text{ mmol}}{1050 \text{ mL}}\right) \left(\frac{6 \text{ mmol}}{1050}\right)^3$

$= 5.33 \times 10^{-10}$

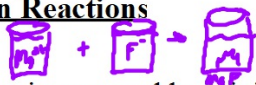
Compare:  $Q > K_{sp}$

$5.33 \times 10^{-10} > 1.9 \times 10^{-10}$

\*  $Q > K_{sp}$  ppt forms

$V_T = 750 + 300 = 1050 \text{ mL}$

Precipitation Reactions



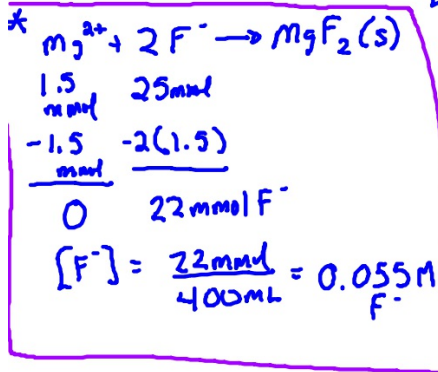
12. A solution is prepared by mixing 150.0 mL of  $1.00 \times 10^{-2} M$   $Mg(NO_3)_2$  and 250.0 mL of  $1.00 \times 10^{-1} M$  NaF. Calculate the concentrations of  $Mg^{2+}$  and  $F^-$  at equilibrium with solid  $MgF_2$ .

Given

$$150.0 \text{ mL} \times 1.00 \times 10^{-2} M Mg^{2+} = 1.5 \text{ mmol } Mg^{2+}$$

$$250.0 \text{ mL} \times 1.00 \times 10^{-1} M F^- = 25 \text{ mmol } F^-$$

$$V_T = 150 + 250 = 400 \text{ mL}$$



Solubility

Topic#8

$$25 \text{ mmol } F^- - 3.0 \text{ mmol} = 22 \text{ mmol } F^- \quad [F^-] = \frac{22 \text{ mmol}}{400 \text{ mL}} = 0.055 M$$

( $K_{sp} = 6.4 \times 10^{-9}$ ). (Ans:  $[Mg^{2+}] = 2.1 \times 10^{-6} M$  and  $[F^-] = 5.50 \times 10^{-2} M$ )

	<u>NTK</u>		
R:	$MgF_2(s) \rightleftharpoons Mg^{2+}(aq) + 2F^-(aq)$		
I:	-	0	0.055
C:	-x	+x	+2x
E:	-	x	<del>2x</del> + 0.055

UNK

$$[Mg^{2+}] = \underline{2.1 \times 10^{-6} M}$$

$$[F^-] = \underline{0.055 M}$$

SOLVE:  $K_{sp} = [Mg^{2+}][F^-]^2$

$$6.4 \times 10^{-9} = (x)(0.055)^2$$

$$x = \frac{6.4 \times 10^{-9}}{(0.055)^2} = 2.1 \times 10^{-6} = [Mg^{2+}]$$

$$[F^-] = 0.055 + 2(2.1 \times 10^{-6}) = 0.0550042$$

**Solubility**  
**Topic#8**

**Precipitation Reactions**

Selective precipitation involves a solution containing two ions. The process involves adding an ion into solution that makes an insoluble ionic compound with BOTH of the ions in solution. The  $K_{sp}$ 's of these insoluble ionic compounds will be magnitudes in difference, so we can predict which will precipitate out first and the concentration of the added ion at the point of precipitation.

13. A solution contains  $1.0 \times 10^{-4} M$   $Cu^+$  and  $2.0 \times 10^{-3} M$   $Pb^{2+}$ . If a source of  $I^-$  is added gradually to this solution, will  $PbI_2$  ( $K_{sp} = 1.4 \times 10^{-8}$ ) or  $CuI$  ( $K_{sp} = 5.3 \times 10^{-12}$ ) precipitate first? Specify the concentration of  $I^-$  necessary to begin precipitation of each salt.

(Ans:  $CuI$  will precipitate first;  $[I^-]$  for  $PbI_2$  is  $2.6 \times 10^{-3} M$  and for  $CuI$  is  $5.3 \times 10^{-8} M$ )

<p><u>Given</u></p> $K_{sp}(PbI_2) = 1.4 \times 10^{-8}$ $K_{sp}(CuI) = 5.3 \times 10^{-12}$ $[Cu^+] = 1.0 \times 10^{-4} M$ $[Pb^{2+}] = 2.0 \times 10^{-3} M$	<p><u>NTK</u></p> $K_{sp} = [Pb^{2+}][I^-]^2$ $1.4 \times 10^{-8} = (2.0 \times 10^{-3})[I^-]^2$ $\sqrt{\frac{1.4 \times 10^{-8}}{2.0 \times 10^{-3}}} = [I^-] = \underline{2.6 \times 10^{-3} M}$ <p style="text-align: right;"><u>unk</u> solve for <math>[I^-]</math></p> $K_{sp} = [Cu^+][I^-]$ $5.3 \times 10^{-12} = (1.0 \times 10^{-4})[I^-]$ $\frac{5.3 \times 10^{-12}}{1.0 \times 10^{-4}} = [I^-] = \underline{5.3 \times 10^{-8} M}$ <p style="text-align: right;"><math>Cu^{2+}</math> ppt out 1<sup>st</sup></p>
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