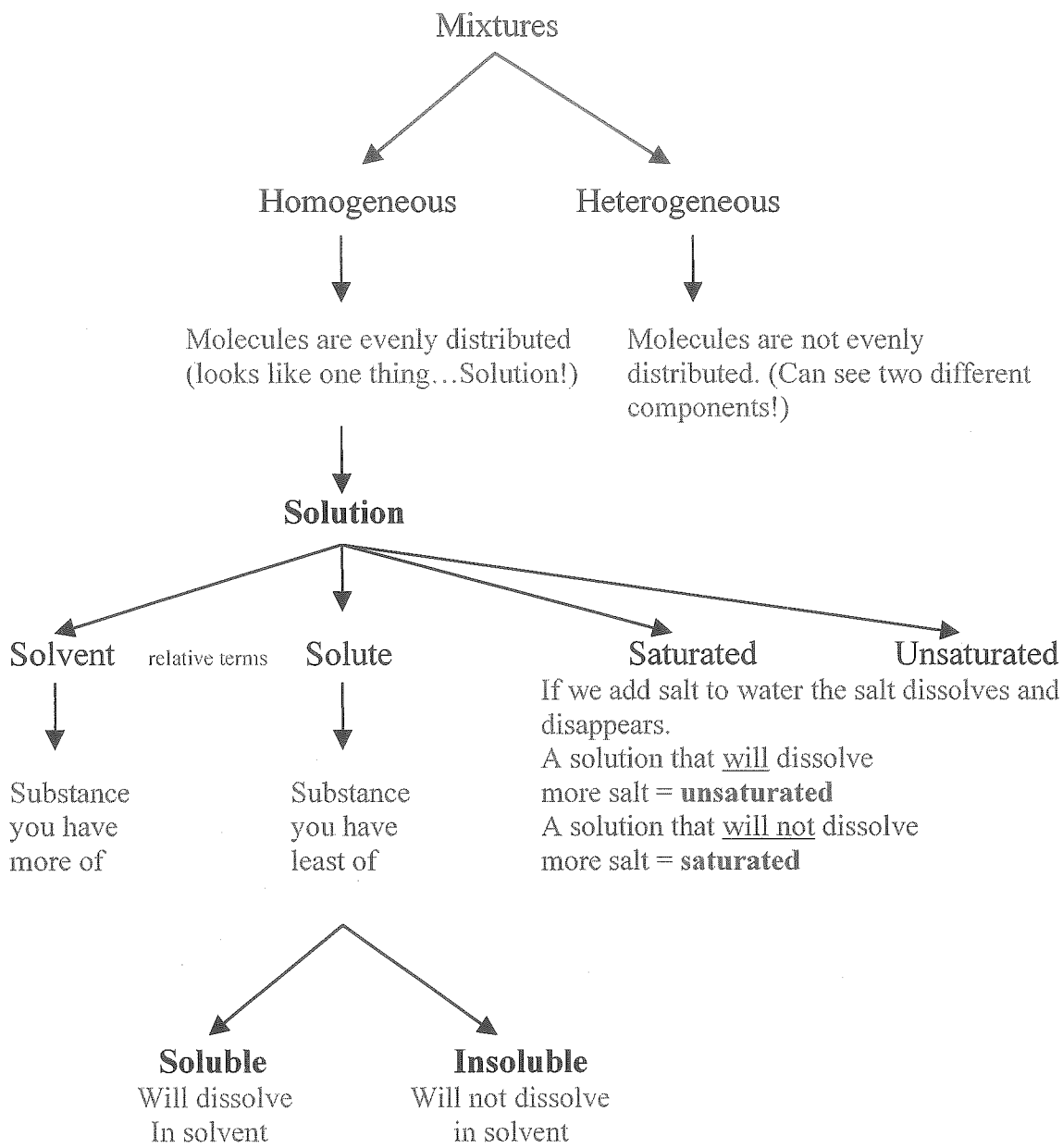


# 1. Solubility Concept Review

## a) Classification



**Solubility** = amount of solute needed to make a saturated solution at a given temperature.

**Molar solubility** = when the amount of solute is expressed as a concentration (mol/L)

## b) The Dissolving Process

i) Do all soluble substances dissolve in the same manner? No

Ionic compounds



Ions



Electrolyte  
(conducts electricity)

Covalent compounds (Molecular)



Neutral molecules



Non-electrolyte  
(non conductor)

ii) Molecular or Ionic solution? How can we tell?

Ionic



Metal (+) and non metal (-)

(polyatomic ions count also!)

Molecular



Non metal and non metal

(organic compounds)

iii) Dissociation

- when an ionic compound is placed in water, the polar water molecules “hydrate” (surround) the ions, pulling them apart.
- thus the ionic compound “dissociates” into its component ions.
- eg:  $\text{NaCl}_{(s)} \xrightarrow{\text{water}} \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)}$
- if undissolved solute exists in the saturated solution, some particles are dissolving and some are going back to the solid form at the same time!!
- a saturated solution is thus in equilibrium!
- eg:  $\text{NaCl}_{(s)} \rightleftharpoons \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)}$
- we can then express molar solubility as the concentration of NaCl required to achieve equilibrium!

## 2. Molar Solubility ← Moles per litre !

*Think of molar solubility as molarity of a saturated solution!!*

Example: 0.49 g of  $\text{AgBrO}_3$  is required to saturate 250 mL of water. What is the molar solubility of  $\text{AgBrO}_3$ ?

$$0.49 \text{ g} \times \frac{1 \text{ mol}}{235.8 \text{ g}} = 0.0021 \text{ mol} \qquad \frac{0.0021 \text{ mol}}{0.25 \text{ L}} = \boxed{0.0083 \text{ M}}$$

Example:  $\text{AlF}_3$  has a molar solubility of 0.0665 M. How many grams of  $\text{AlF}_3$  will dissolve in 1 L?

$$\text{Mol} = \text{M} \times \text{Litre} = 0.0665 \text{ M} \times 1 \text{ L} = 0.0665 \text{ mol AlF}_3$$

$$0.0665 \text{ mol} \times \frac{84 \text{ g}}{1 \text{ mol}} = \boxed{5.59 \text{ g / 1 L}}$$

## 3. Dilution Calculations

*Use the formula:*

$$\mathbf{C_1 V_1 = C_2 V_2}$$

(initial conc.) x (initial volume) = (final conc.) x (final volume)

Example: Suppose you have 300.0 mL of 1.2 M HCl. How much water do you need to add to dilute it to 0.50 M HCl?

$$C_1 V_1 = C_2 V_2$$

$$(1.2 \text{ M})(300 \text{ mL}) = (0.5 \text{ M})(V_2)$$

$$V_2 = 720 \text{ mL}$$

$$720 \text{ mL} - 300 \text{ mL} = \boxed{420 \text{ mL water added}}$$

Example: You have 145 mL of a 8.0 M  $\text{HNO}_3$  solution. What is final concentration after adding 650 mL  $\text{H}_2\text{O}$ ?

$$C_1 V_1 = C_2 V_2$$

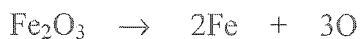
$$(8 \text{ M})(145 \text{ mL}) = (C_2)(145 + 650 \text{ mL})$$

$$C_2 = 1.46 \text{ M} \sim \boxed{1.5 \text{ M}}$$

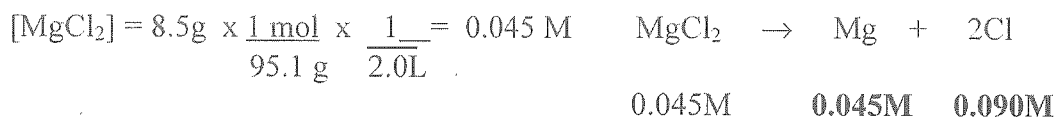
## 4. Concentration of Individual Ions



Example: What are the concentrations of ions in a 1.5 M  $\text{Fe}_2\text{O}_3$  solution?



Example: 8.5g of  $\text{MgCl}_2$  is dissolved in 2.0L. What is  $[\text{Mg}^{+2}]$  and  $[\text{Cl}^-]$ ?



Example: 500 mL of 0.8 M  $\text{Li}_2\text{CO}_3$  is mixed with 500 mL of 0.5 M  $\text{BeF}_2$ . Find concentration of all ions!



$$C_1V_1 = C_2V_2$$

$$(0.8)(0.5) = (C_2)(1.0)$$

$$C_2 = 0.40 \text{ M}$$

$$0.80\text{M} \quad 0.40\text{M}$$



$$C_1V_1 = C_2V_2$$

$$(0.5)(0.5) = (C_2)(1.0)$$

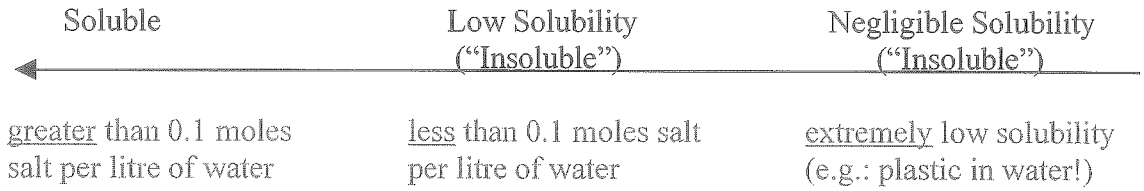
$$C_2 = 0.25 \text{ M}$$

$$0.25\text{M} \quad 0.50\text{M}$$

(III. 3 + 4)

## 5. Solubility of Ionic Compounds (Salts)

### a) Solubility Range of Salts in Water



### b) Predicting Solubility

i) Must Use:

Table 16.3 page 459 in Heath *or*  
Table "Solubility of Common Compounds" page 332 in Hebden

ii) Example: Is  $\text{FeCl}_3$  soluble in water?

Look in negative ion column and see that  $\text{Cl}^-$  is soluble with any positive ion except  $\text{Ag}^+$ ,  $\text{Pb}^{+2}$ ,  $\text{Hg}^{+2}$  or  $\text{Cu}^+$

Answer: SOLUBLE

iii) Example: Is  $\text{MgSO}_4$  soluble in water?

Answer: SOLUBLE

iv) Example: Is  $\text{PbS}_2$  soluble in water?

Answer: INSOLUBLE

### c) Precipitate

i) What is a "precipitate"?

A solid formed when two liquids or aqueous solutions react.

ii) Predicting a precipitate

"Mix and match" the ions from the aqueous solutions and refer to a solubility table to determine if any or all ionic compounds are soluble.

iii) Example: Will a precipitate form from mixing  $\text{BaS}_{(aq)}$  and  $\text{AgNO}_3_{(aq)}$ ?

① Write reaction:  $\text{BaS} + 2\text{AgNO}_3 \rightleftharpoons \text{Ba(NO}_3)_2 + \text{Ag}_2\text{S}$

② Ignore  $\text{BaS}$  and  $\text{AgNO}_3$  because if we have solution of these they must be soluble!  
(You can check solubility table if you want but you waste your time!)

③ Check  $\text{Ba(NO}_3)_2$ .....SOLUBLE  
Check  $\text{Ag}_2\text{S}$  .....INSOLUBLE (Low Solubility)

④ You will form a precipitate of  $\text{Ag}_2\text{S}_{(s)}$

iv) Example: Will a precipitate form from mixing  $\text{KI}_{(aq)}$  and  $\text{Pb(NO}_3)_2_{(aq)}$ ?

① Write reaction:  $2\text{KI} + \text{Pb(NO}_3)_2 \rightleftharpoons 2\text{KNO}_3 + \text{PbI}_2$

② Ignore  $\text{KI}$  and  $\text{Pb(NO}_3)_2$  because if we have solution of these they must be soluble!

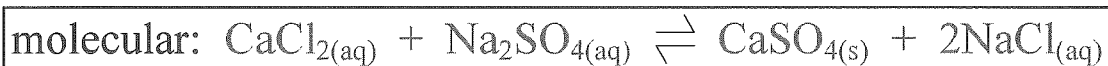
③ Check  $\text{KNO}_3$ .....SOLUBLE  
Check  $\text{PbI}_2$  .....INSOLUBLE (Low Solubility)

④ You will form a precipitate of  $\text{PbI}_2_{(s)}$

#### d) Writing Equations

i) This is a review of “equilibrium systems in solutions” from Unit II

ii) Review Ionic Equations:



iii)  $\text{Na}^+$  and  $\text{Cl}^-$  appear on both sides of the equilibrium so they cancel out. They are called spectator ions.

Do Questions: # 21-23 page 83; #24 page 84; #25 page 87

## 6. Using Solubility Table and Precipitation to Identify Ions

### a) Where would a chemist want to do this?

We might have a solution containing unknown ions and wish to determine what ions are present.

### b) What is the process used to identify ions?

i) Lets pretend we suspect a solution contains  $\text{Ba}^{+2}$  or  $\text{Cu}^+$ .

ii) Search the Solubility Table for the anions that can precipitate (ppt.) one or both cations.

We quickly find that  $\text{SO}_4^{-2}$ ,  $\text{OH}^-$ ,  $\text{PO}_4^{-3}$ ,  $\text{CO}_3^{-2}$ ,  $\text{SO}_3^{-2}$  will precipitate  $\text{Ba}^{+2}$ .

We also find that  $\text{Cl}^-$ ,  $\text{I}^-$ ,  $\text{Br}^-$ ,  $\text{S}^{-2}$ ,  $\text{OH}^-$ ,  $\text{PO}_4^{-3}$ ,  $\text{CO}_3^{-2}$ ,  $\text{SO}_3^{-2}$  will precipitate  $\text{Cu}^+$

iii) Make a table to organize your ions:

	$\text{Cl}^-$ , $\text{I}^-$ , $\text{Br}^-$	$\text{SO}_4^{-2}$	$\text{S}^{-2}$	$\text{OH}^-$	$\text{PO}_4^{-3}$ , $\text{CO}_3^{-2}$ , $\text{SO}_3^{-2}$
$\text{Cu}^+$	ppt	--	ppt	ppt	ppt
$\text{Ba}^{+2}$	--	ppt	--	ppt	ppt

iv) Conclude that if we add  $\text{Cl}^-$  (or  $\text{I}^-$ ,  $\text{Br}^-$ ) or  $\text{S}^{-2}$ , and a ppt. forms,  $\text{Cu}^+$  is present. If we add  $\text{SO}_4^{-2}$  and a ppt. forms,  $\text{Ba}^{+2}$  is present.

### c) What would be the experimental procedure used to identify the ions?

**Step 1:** To 1ml unknown solution, add a few drops of 1M HCl. (or sulphide salt!)

If a ppt. forms,  $\text{Cu}^+$  is present. Filter and discard ppt. and proceed to Step 2 with the remaining solution.

If a ppt. does not form.  $\text{Cu}^+$  is not present. Proceed to Step 2.

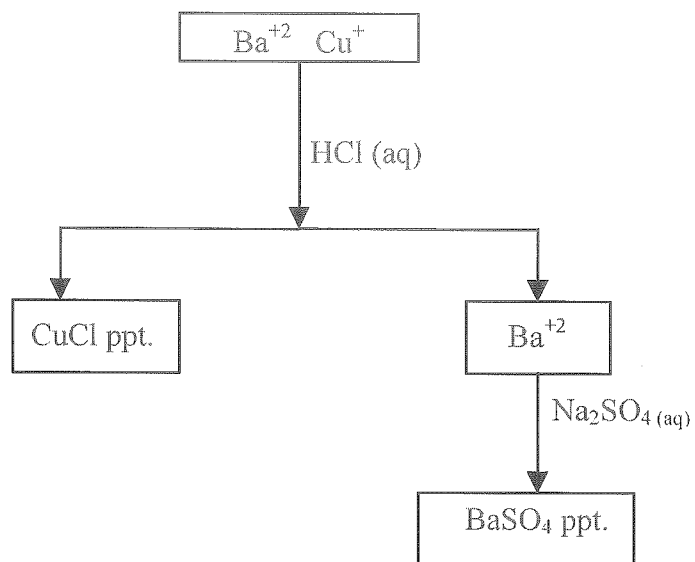
**Step 2:** To the solution from Step 1, add a few drops of 1M  $\text{Na}_2\text{SO}_4$ .

If a ppt. forms,  $\text{Ba}^{+2}$  is present.

If a ppt. does not form.  $\text{Ba}^{+2}$  is not present.

**d) Flow Charts Rule!**

Use flow charts to help you organize the proper experimental procedure!



Hw: Read through examples on page 88-89, section III.5  
Do questions: #26-37 page 90-91



(III, 6) 6  
Day 1  
PS #4

## 7. Ksp and Solubility

In the last sections, we were able to qualitatively identify which salts were soluble or not soluble.

What about salts that are slightly soluble?

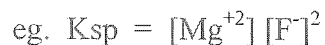
Now, we want to quantify (attach a numerical value) the degree of solubility of the slightly soluble salts.

### a) The Solubility Product

i) Salts that are only slightly soluble will form an equilibrium when they dissolve:

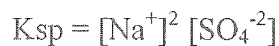


ii) We can write an equilibrium expression for the solubility of a salt.



Note: ① attach "sp" to K symbol when dealing with solubility equilibrium.  
② It works just like equilibrium expressions! Notice how  $[\text{MgF}_2]$  is not included because it is a solid!

iii) Example: Write the Ksp expression for  $\text{Na}_2\text{SO}_{4(s)}$



### b) Meaning of Ksp

i) High Ksp value = higher concentration of ions in solution = High Solubility

ii) Low Ksp value = lower concentration of ions in solution = Low Solubility

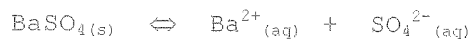
iii) Check out page 333 in Hebden or Appendix B page 722 in Heath

iv) Ksp is a constant.

✓ Example 1 :

The solubility of  $\text{BaSO}_4$  is found to be  $1.05 \times 10^{-5} \text{ M}$  at  $25^\circ\text{C}$ . Calculate the  $K_{\text{sp}}$  for  $\text{BaSO}_4$ .

Answer 1 :



$$[\text{Ba}^{2+}] = 1.05 \times 10^{-5} \text{ M}$$

$$[\text{SO}_4^{2-}] = 1.05 \times 10^{-5} \text{ M}$$

$$K_{\text{sp}} = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$$

$$= (1.05 \times 10^{-5} \text{ M})(1.05 \times 10^{-5} \text{ M})$$

$$= 1.10 \times 10^{-10}$$

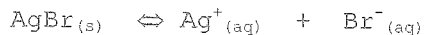
Note : as with  $K_{\text{eq}}$  we don't place units in  $K_{\text{sp}}$ .

✓ Example 2 :

The solubility of  $\text{AgBr}$  is  $1.33 \times 10^{-4} \text{ g/L}$ . Calculate the  $K_{\text{sp}}$  for  $\text{AgBr}$ .

Answer 2 :

$$\begin{aligned} \text{concentration AgBr solution} &= \frac{1.33 \times 10^{-4} \text{ g/L}}{187.8 \text{ g/mol}} \\ &= 7.08 \times 10^{-7} \text{ M} \end{aligned}$$



$$[\text{Ag}^{+}] = [\text{Br}^{-}] = 7.08 \times 10^{-7} \text{ M}$$

$$K_{\text{sp}} = [\text{Ag}^{+}][\text{Br}^{-}]$$

$$= (7.08 \times 10^{-7} \text{ M})(7.08 \times 10^{-7} \text{ M})$$

$$= 5.02 \times 10^{-13}$$

In every  $K_{\text{sp}}$  question, it is recommended that one shows both the dissociation equation and the  $K_{\text{sp}}$  expression. This will clearly show that the student displays the chemical principles involved in the calculations.

~~Period 31~~

~~go over ionic reaction worksheet~~

~~quiz tomorrow on ionic reactions~~

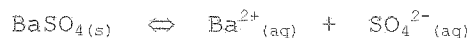
- Solubility

Solubility is defined as how much of a substance we can dissolve in a given amount of solvent. We usually measure this quantity in moles/litre or grams/litre.

To determine the solubility, we take a given amount of a saturated solution and evaporate the water to determine the mass of solute present. It is important to use a saturated solution because this shows that the solubility point has been reached.

- Solubility Equilibria (Ksp)

In lab 16-E we saw that a precipitate was formed when  $Ba^{2+}$  ions were mixed with  $SO_4^{2-}$  ions. According to our table of solubilities, this was to be expected. However, this does not mean that NO  $BaSO_4$  could be dissolved in water. In fact, there was some there, albeit a small amount. If we were to take some solid  $BaSO_4$  and try to dissolve it in distilled water, some would dissolve but the majority of the solid would remain as such in the solution. We reach the solubility point rather quickly and in effect set up an equilibrium :



Because we can set up an equilibrium, we can also write an equilibrium constant expression for this reaction :

$$K_{eq} = [Ba^{2+}][SO_4^{2-}]$$

This particular equilibrium constant occurs so frequently, we give it its own name; the solubility products constant and its own symbol Ksp

$$K_{sp} = [Ba^{2+}][SO_4^{2-}]$$

We can easily calculate the Ksp of a substance from the solubility data.

### c) Experimentally Finding K<sub>sp</sub>

**Method 1:** ① Simply take MgF<sub>2(s)</sub> and add to water until solution is saturated.

② If we know mass of MgF<sub>2</sub> added and water volume we can find [MgF<sub>2</sub>] and then we know that:

$$[\text{Mg}^{+2}] = [\text{MgF}_2] \quad \text{and} \quad [\text{F}^-] = 2 \times [\text{MgF}_2]$$

③  $K_{sp} = [\text{Mg}^{+2}] [\text{F}^-]^2$

**Method 2:** ① Mix together a source of Mg<sup>+2</sup> such as MgSO<sub>4(aq)</sub> and a source of F<sup>-</sup> such as NaF<sub>(aq)</sub>.

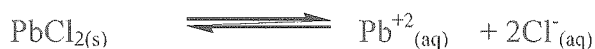
② Let ppt. of MgF<sub>2(s)</sub> form and “analyze” solution to find [Mg<sup>+2</sup>] and [F<sup>-</sup>].

③ It is difficult to find [Mg<sup>+2</sup>] and [F<sup>-</sup>] but there *are* analytical methods!

### d) K<sub>sp</sub> Calculations

#### Type 1 (Find K<sub>sp</sub> from ion concentrations)

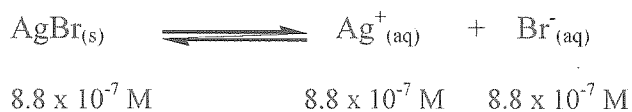
What is the K<sub>sp</sub> for PbCl<sub>2</sub> if [Pb<sup>+2</sup>] is 1.1 x 10<sup>-4</sup> M and [Cl<sup>-</sup>] is 0.33 M?



$$K_{sp} = [\text{Pb}^{+2}] [\text{Cl}^-]^2 \qquad K_{sp} = (1.1 \times 10^{-4} \text{ M})(0.33 \text{ M})^2$$

$$K_{sp} = 1.2 \times 10^{-5}$$

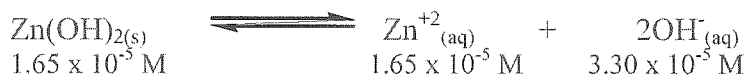
What is the K<sub>sp</sub> for AgBr if the solubility of AgBr is 8.8 x 10<sup>-7</sup> M?



$$K_{sp} = [\text{Ag}^+] [\text{Br}^-] = (8.8 \times 10^{-7} \text{ M})^2 = 7.7 \times 10^{-13}$$

If 1.64 x 10<sup>-6</sup> g of Zn(OH)<sub>2</sub> can dissolve in 1.0 mL of water, what is the K<sub>sp</sub>?

$$(1.64 \times 10^{-6} \text{ g}/1.0 \text{ mL}) \times (1 \text{ mol}/99.38 \text{ g}) \times (1000 \text{ mL}/1 \text{ L}) = 1.65 \times 10^{-5} \text{ M}$$



$$K_{sp} = [\text{Zn}^{+2}][\text{OH}^-]^2 = (1.65 \times 10^{-5} \text{ M})(3.30 \times 10^{-5} \text{ M})^2 = 1.8 \times 10^{-14}$$

## Example 3 :

The solubility of  $\text{Ag}_2\text{S}$  is  $3.2 \times 10^{-15}$  g/L at  $25^\circ\text{C}$ . Calculate the  $K_{\text{sp}}$ .

Answer 3 :

$$\text{concentration } \text{Ag}_2\text{S} = \frac{3.2 \times 10^{-15} \text{ g/L}}{247.9 \text{ g/mol}} = 1.3 \times 10^{-17} \text{ M}$$



$$[\text{Ag}^+] = (2)(1.3 \times 10^{-17} \text{ M}) = 2.6 \times 10^{-17} \text{ M}$$

$$[\text{S}^{2-}] = 1.3 \times 10^{-17} \text{ M}$$

$$K_{\text{sp}} = [\text{Ag}^+]_2[\text{S}^{2-}]$$

$$= (2.6 \times 10^{-17} \text{ M})^2(1.3 \times 10^{-17} \text{ M})$$

$$= 8.6 \times 10^{-51}$$

- Assignment :

Page 548 #1,3

- Quiz next day on ionic reaction worksheet

Period 32

- Quiz on ionic reaction worksheet
- Go over questions Page 548 #1, 3
- Ksp to solubility PS#5

Example 4 :

The Ksp of AgI is  $8.3 \times 10^{-17}$ . Calculate its solubility in mol/L.

Answer 4 :

Set the answer up like a keq ICE question.

	$\text{AgI}_{(s)}$	$\Leftrightarrow$	$\text{Ag}^+_{(aq)}$	+	$\text{I}^-_{(aq)}$
I	(X)		0		0
C			+X		+X
E			X		X

$$K_{sp} = [\text{Ag}^+][\text{I}^-]$$

$$8.3 \times 10^{-17} = (X)(X)$$

$$8.3 \times 10^{-17} = X^2$$

$$9.1 \times 10^{-9} = X$$

If the solution is set up in this form, X will always turn out to be the solubility of the compound in moles per litre.

$9.1 \times 10^{-9}$  M is the solubility of AgI.

**Type 2 (Find ion concentrations from Ksp value)**

What is the concentration of  $\text{Ca}^{+2}$  and  $\text{CO}_3^{-2}$  ions if the Ksp for  $\text{CaCO}_3$  is  $4.8 \times 10^{-9}$ ?



② Let  $[\text{CaCO}_3] = X$



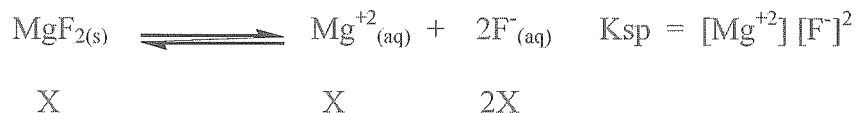
④  $K_{sp} = [\text{Ca}^{+2}][\text{CO}_3^{-2}]$

⑤  $4.8 \times 10^{-9} = [X][X] \quad 4.8 \times 10^{-9} = X^2$

$X = 6.9 \times 10^{-5} \text{ M}$  Therefore;  $[\text{Ca}^{+2}] = [\text{CO}_3^{-2}] = 6.9 \times 10^{-5} \text{ M}$

The Ksp for  $\text{MgF}_2$  is  $6.4 \times 10^{-9}$ .

a) What is the  $[\text{Mg}^{+2}]$  and  $[\text{F}^-]$ ?



$6.4 \times 10^{-9} = [\text{Mg}^{+2}][\text{F}^-]^2$

$6.4 \times 10^{-9} = [X][2X]^2$

$6.4 \times 10^{-9} = 4X^3$

$1.2 \times 10^{-3} \text{ M} = X$  Therefore;  $[\text{Mg}^{+2}] = 1.2 \times 10^{-3} \text{ M}$  and  $[\text{F}^-] = 2.4 \times 10^{-3} \text{ M}$

b) What is the molar solubility of  $\text{MgF}_2$ ?

Molar solubility is mol per litre or M.

Therefore, molar solubility of  $\text{MgF}_2$  is simply  $1.2 \times 10^{-3} \text{ M}$

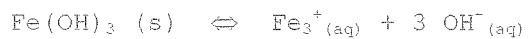
c) What is the solubility of  $\text{MgF}_2$  in g/L?

$1.2 \times 10^{-3} \text{ mol/L} \times 62.3 \text{ g/mol} = 7.3 \times 10^{-2} \text{ g/L}$

Example 5 :

The  $K_{sp}$  of  $\text{Fe}(\text{OH})_3$  is  $1.1 \times 10^{-36}$  at  $18^\circ\text{C}$ . Calculate the solubility in g/L.

Answer 5 :



I	(X)	0	0
C		+X	+3X
E		X	3X

$$K_{sp} = [\text{Fe}^{3+}][\text{OH}^{-}]^3$$

$$1.1 \times 10^{-36} = (X)(3X)^3$$

$$1.1 \times 10^{-36} = 27X^4$$

$$4.1 \times 10^{-38} = X^4$$

$$4.5 \times 10^{-10} = X$$

$$(4.5 \times 10^{-10} \text{ mol/L})(106.8 \text{ g/mol}) = 4.8 \times 10^{-8} \text{ g/L}$$

- Assignment :

Page 548 #2 a,c  
#4 a,b  
Solubility worksheet



## 8. K<sub>sp</sub> and Precipitate Formation

The value of K<sub>sp</sub> is the point between the ions precipitating or the ions remaining soluble.

At exactly K<sub>sp</sub>, the ions will *just* start to precipitate. You have just arrived at the saturation point.

### a) "Trial K<sub>sp</sub>" aka. "The Ion Product" aka. "Q"

- i) this is a K<sub>sp</sub> value for the ion concentrations before mixing with other ion(s).
- ii) we will have to calculate a Trial K<sub>sp</sub> value using the initial ion concentrations, to determine if mixing will result in a precipitate.

**Q < K<sub>sp</sub>** Then the ions will stay in solution.  
Not enough ions to precipitate.

**Q > K<sub>sp</sub>** Then the solution cannot dissolve more ions.  
The undissolved ions will join and precipitate.

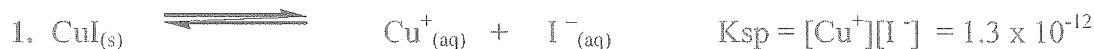
**Q = K<sub>sp</sub>** There are just enough ions to barely form a saturated solution.  
You are at the saturation point.

### b) Predicting Precipitate Calculations

#### Type 1 (will the ions precipitate when mixed together?)

- General Strategy:
1. Write equilibrium equation and find K<sub>sp</sub>
  2. Find ion concentration
  3. Find Q
  4. Compare Q with K<sub>sp</sub>

**Example:** Will a precipitate form when 4.5 mL of  $5.0 \times 10^{-4}$  M  $\text{Cu}^+$  is mixed with 2.0 mL of  $6.0 \times 10^{-5}$  M  $\text{I}^-$ ?



2. Find ion concentrations:

Mixing the two solutions results in a dilution, so we must recalculate the concentrations.

$$[\text{Cu}^+] \quad C_1V_1 = C_2V_2 \quad (5.0 \times 10^{-4} \text{ M})(4.5\text{mL}) = C_2(6.5\text{mL}); [\text{Cu}^+] = 3.5 \times 10^{-4} \text{ M}$$

$$[\text{I}^-] \quad C_1V_1 = C_2V_2 \quad (6.0 \times 10^{-5} \text{ M})(2.0\text{mL}) = C_2(6.5\text{mL}); [\text{I}^-] = 1.8 \times 10^{-5} \text{ M}$$

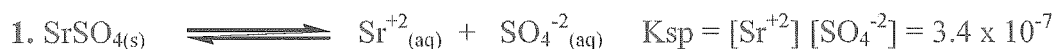
3. Find Q:

$$Q = [\text{Cu}^+][\text{I}^-] = [3.5 \times 10^{-4} \text{ M}][1.8 \times 10^{-5} \text{ M}] = 6.0 \times 10^{-9}$$

4. Compare to  $K_{sp}$ :

$$\begin{array}{ccc} Q & > & K_{sp} \\ (6.0 \times 10^{-9}) & & (1.3 \times 10^{-12}) \end{array} \quad \text{Thus, a precipitate will form!}$$

**Example:** Will a precipitate form when 12.0 mL of  $1.3 \times 10^{-3}$  M  $\text{Sr}^{+2}$  is mixed with 4.0 mL  $2.0 \times 10^{-9}$  M  $\text{SO}_4^{-2}$ ?



$$2. [\text{Sr}^{+2}] = (12.0 \text{ mL})(1.3 \times 10^{-3}) / 16.0 \text{ mL} = 9.8 \times 10^{-4} \text{ M}$$

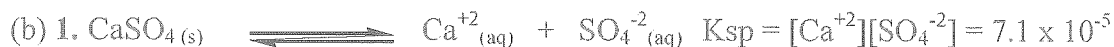
$$[\text{SO}_4^{-2}] = (4.0 \text{ mL})(2.0 \times 10^{-9}) / 16.0 \text{ mL} = 5.0 \times 10^{-10} \text{ M}$$

$$3. Q = [\text{Sr}^{+2}][\text{SO}_4^{-2}] = (9.8 \times 10^{-4} \text{ M})(5.0 \times 10^{-10} \text{ M}) = 4.9 \times 10^{-13}$$

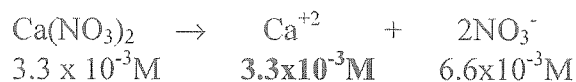
$$\begin{array}{ccc} K_{sp} & > & Q \\ (3.4 \times 10^{-7}) & & (4.9 \times 10^{-13}) \end{array} \quad \text{Thus, a precipitate will NOT form!}$$

**Example:** If 20.0 mL of  $5.0 \times 10^{-3}$  M  $\text{Ca}(\text{NO}_3)_2$  is mixed with 10.0 mL of  $2.5 \times 10^{-2}$  M  $\text{Sn}(\text{SO}_4)_2$ : (a) What precipitate is possible?  
 (b) Will this precipitate form at the given concentrations?

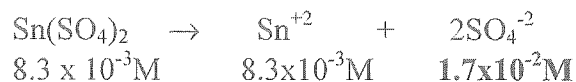
(a)  $\text{CaSO}_4$  is the only combo of ions that will potentially precipitate according to the solubility table. ( $\text{Sn}(\text{NO}_3)_4$  is soluble at all concentrations)



2.  $[\text{Ca}(\text{NO}_3)_2] = (20.0 \text{ mL})(5.0 \times 10^{-3}) / 30.0 \text{ mL} = 3.3 \times 10^{-3} \text{ M}$



$[\text{Sn}(\text{SO}_4)_2] = (10.0 \text{ mL})(2.5 \times 10^{-2}) / 30.0 \text{ mL} = 8.3 \times 10^{-3} \text{ M}$



3.  $Q = [\text{Ca}^{+2}][\text{SO}_4^{-2}] = (3.3 \times 10^{-3} \text{ M})(1.7 \times 10^{-2} \text{ M}) = 5.6 \times 10^{-5}$

4.  $K_{\text{sp}} > Q$  Thus, a precipitate will NOT form!  
 $(7.1 \times 10^{-5}) > (5.6 \times 10^{-5})$

### Type 2 (what [ion] is required to precipitate out the other ion?)

- General Strategy:
1. Write equilibrium equation and find  $K_{\text{sp}}$
  2. Know  $Q = K_{\text{sp}}$  at this particular [ion]
  3. Solve  $K_{\text{sp}}$  equation for other [ion]

**Example:** If we have a  $3.0 \times 10^{-4}$  M solution of  $\text{Pb}^{+2}$ , what  $[\text{I}^-]$  is required to just start the precipitation of  $\text{PbI}_2(\text{s})$  from the solution?



2.  $Q = K_{\text{sp}} = 8.5 \times 10^{-9}$  (this means we can use the  $[\text{Pb}^{+2}]$  given)

3.  $[\text{I}^-]^2 = K_{\text{sp}} / [\text{Pb}^{+2}] = 8.5 \times 10^{-9} / 3.0 \times 10^{-4} = 2.8 \times 10^{-5}$

$[\text{I}^-] = \sqrt{(2.8 \times 10^{-5})} = 5.3 \times 10^{-3} \text{ M}$

Period 34

- Quiz on solubility
- Go over questions Page 554 #5-8
- Predicting when precipitation occurs (Version 2.0)

7.6 p. 487-190

We can use the  $K_{sp}$  value to predict whether or not a precipitate will form in a given solution provided that we know the concentration of the ions that appear in the  $K_{sp}$  expression. For example, we know that in a saturated solution of  $\text{CaCO}_3$ , the ion product  $[\text{Ca}^{2+}][\text{CO}_3^{2-}]$  is EXACTLY equal to  $k_{sp}$ . In an unsaturated solution, the concentrations of  $\text{Ca}^{2+}$  and  $\text{CO}_3^{2-}$  are less than those in a saturated solution therefore their product is less than  $K_{sp}$ . No precipitate will form in an unsaturated solution. If, however, the ion product is greater than the  $K_{sp}$  value, then there will be a precipitate formed.

Summary :

If ion product ( $K_{trial}$ )  $>$   $K_{sp}$  there will be a ppt formed.

If  $K_{trial} = K_{sp}$  we have a saturated solution. No ppt will form.

If  $K_{trial} < K_{sp}$  we have an unsaturated solution. No ppt will form.

Example 1 :

A student wishes to prepare 1.0 L of a solution containing 0.015 mol NaCl and 0.15 mol of  $\text{Pb}(\text{NO}_3)_2$ . She was concerned that a precipitate of  $\text{PbCl}_2$  might form. The  $K_{sp}$  of  $\text{PbCl}_2$  is  $1.7 \times 10^{-5}$ . Can she expect a precipitate of  $\text{PbCl}_2$  to form?

Answer 1 :

$$K_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]_2$$

$$[\text{Pb}^{2+}] = 0.15 \text{ mol}/1.0 \text{ L} = 0.15 \text{ M}$$

$$[\text{Cl}^-] = 0.015 \text{ mol}/1.0 \text{ L} = 0.015 \text{ M}$$

$$\begin{aligned} K_{trial} &= (0.15 \text{ M})(0.015 \text{ M})^2 \\ &= 3.4 \times 10^{-5} \end{aligned}$$

Because  $K_{trial} (3.4 \times 10^{-5}) > K_{sp} (1.7 \times 10^{-5})$ , a ppt will form.

Example 2 :

Will a ppt of  $\text{PbSO}_4$  form if 100.0 mL of  $1.0 \times 10^{-3} \text{ M Pb(NO}_3)_2$  is mixed with 100.0 mL of  $2.0 \times 10^{-3} \text{ M MgSO}_4$  ?  $K_{sp} \text{ PbSO}_4 = 6.3 \times 10^{-7}$

Answer 2 :

$$[\text{Pb(NO}_3)_2] = (1.0 \times 10^{-3} \text{ M}) \frac{100.0 \text{ mL}}{200.0 \text{ mL}} = 5.0 \times 10^{-4} \text{ M}$$

$$[\text{MgSO}_4] = (2.0 \times 10^{-3} \text{ M}) \frac{100.0 \text{ mL}}{200.0 \text{ mL}} = 1.0 \times 10^{-3} \text{ M}$$

$$[\text{Pb}^{2+}] = 5.0 \times 10^{-4} \text{ M}$$

$$[\text{SO}_4^{2-}] = 1.0 \times 10^{-3} \text{ M}$$

$$K_{sp} = [\text{Pb}^{2+}][\text{SO}_4^{2-}]$$

$$K_{\text{trial}} = (5.0 \times 10^{-4} \text{ M})(1.0 \times 10^{-3} \text{ M})$$

$$= 5.0 \times 10^{-7}$$

$K_{\text{trial}} < K_{sp}$  therefore no precipitate will form.

Example 3 :

Will a precipitate of  $\text{PbCl}_2$  form when 50.0 mL of 0.10 M  $\text{Pb(NO}_3)_2$  is added to 20.0 mL of 0.040 M NaCl ?  
 $K_{sp} \text{ PbCl}_2 = 1.7 \times 10^{-5}$

Answer 3 :

$$[\text{Pb(NO}_3)_2] = (0.10 \text{ M}) \frac{50.0 \text{ mL}}{70.0 \text{ mL}} = 0.071 \text{ M}$$

$$[\text{NaCl}] = (0.040 \text{ M}) \frac{20.0 \text{ mL}}{70.0 \text{ mL}} = 0.011 \text{ M}$$

$$[\text{Pb}^{2+}] = 0.071 \text{ M}$$

$$[\text{Cl}^-] = 0.011 \text{ M}$$

$$K_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]_2$$

$$K_{\text{trial}} = (0.071 \text{ M})(0.011 \text{ M})^2$$

$$= 9.3 \times 10^{-6}$$

$K_{\text{trial}} > K_{sp}$  therefore a precipitate will form.

- Assignment :

Page 554 #1,2  
 Precipitation Worksheet

- Set up lab 19-C

## 9. Solubility and Le Chatelier's Effect



recall solubility  
equilibrium  
if we shift to  
the right, there is  
more product

### a) Shifting Equilibrium to Left

- i) If shift the equilibrium to the left, we are *decreasing* the solubility of the  $\text{CuI}_{(s)}$  salt.
- ii) We can shift the equilibrium to left by increasing  $[\text{Cu}^+]$  or  $[\text{I}^-]$
- iii) We must add only  $\text{Cu}^+$  or  $\text{I}^-$ . If we add both, they will just combine to form  $\text{CuI}_{(s)}$ !
- iv) So we add a salt containing one ion in common with the equilibrium.  
Eg.  $\text{CuNO}_3$  or  $\text{NaI}$
- v) This procedure is called the "Common Ion Effect".

### a) Shifting Equilibrium to Right

- i) We are *increasing* the solubility of the  $\text{CuI}$  salt.
- ii) Recall, in equilibrium expressions, changing a solid's concentration will not affect the equilibrium. So cannot add more  $\text{CuI}_{(s)}$ !
- iii) Therefore, we must decrease the  $[\text{Cu}^+]$  or  $[\text{I}^-]$  to shift the equilibrium to right.
- iv) General technique is to add an ion that will precipitate out some  $\text{Cu}^+$  or  $\text{I}^-$ .  
(Choose an appropriate ion from solubility table)

Eg. Add  $\text{Pb}^{+2}$  (or  $\text{Ag}^+$ ) to form insoluble  $\text{PbI}_{2(s)}$  to remove  $\text{I}^-$ .

Eg. Add  $\text{S}^{-2}$  (or  $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{OH}^-$ ,  $\text{PO}_4^{-3}$ ,  $\text{CO}_3^{-2}$ ,  $\text{SO}_3^{-2}$ ) to precipitate out the  $\text{Cu}^+$ .

c) no shift -  $\text{H}_2\text{O}$ , (s)

Hw: Read through examples on pages 105-107

Do Questions: #81-86 page 108

Period 33

- Go over homework

Page 548 #2 a,c  
#4 a,b  
Solubility worksheet

- Quiz tomorrow on solubility

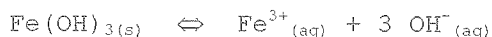
- How can we change the solubility?

PS#8

1. Increase temperature. Because the process of dissolving is usually endothermic, if we increase the temperature of a solution, we can increase the amount of the solute we can dissolve. However, there are some weird compounds that will show a decrease in solubility when the temperature is increased. For this reason, all the solubilities and solubility products are taken at certain temperatures.

2. The 'common ion effect'.

Consider the following equilibrium :



Q. What would happen to the equilibrium if we added some  $\text{Fe}^{3+}$  ions ?

A. Shift left producing more  $\text{Fe}(\text{OH})_{3(s)}$ . This effectively DECREASES the solubility. This is because we can not dissolve as much solid. The same thing would happen if we were to try and dissolve the solid in a solution containing  $\text{Fe}^{3+}$  ions already.

Q. What would happen to the equilibrium if we added  $\text{OH}^{-}$  ions ?

A. Solubility decreases for the same reason as when we added  $\text{Fe}^{3+}$  ions. This decrease in solubility due to an ion that is part of the equilibrium is known as the COMMON ION EFFECT.

P.490-492

Q. What would happen to the equilibrium if we added  $\text{Fe}(\text{OH})_{3(s)}$  ?

A. Nothing. We would just have more solid at the bottom of the beaker. Remember, an equilibrium will only shift if the concentration of a species is increased or decreased. Adding more solid does not change the solid's concentration, just the amount we have.

Q. How then can we increase the solubility of this compound (that is without changing the temperature) ?

A. Add nitric acid. The  $H^+$  ions in the acid would react with the  $OH^-$  ions thus lowering the concentration of the  $OH^-$  ions. This would shift the equilibrium to the right increasing the solubility of  $Fe(OH)_3$ . We add nitric acid because the nitrate ion will not react with the  $Fe^{3+}$  ion therefore no other precipitate will form. If another ppt was to be formed, the solubility of the  $Fe(OH)_3$  would indeed increase but we might not notice it. For example,  $FePO_4$  would precipitate if we added  $H_3PO_4$  but we would not notice a change.  $FePO_4$  looks a lot like  $Fe(OH)_3$ .

PS#9

- Maximum amount of an ion that can be dissolved in a solution.

As we have seen, the amount of a substance that can be dissolved in a solution decreases when there is one of the ions already dissolved in solution. Using the  $K_{sp}$  value, we can determine the maximum amount of a substance that can be added to a solution to not allow a precipitate to form.

Example 1 :

What is the maximum mass of  $Na_2SO_4$  that can be dissolved in 1.00 L of 0.0500 M  $Pb(NO_3)_2$  ?  
 $K_{sp} PbSO_4 = 6.3 \times 10^{-7}$

Answer 1 :

$$K_{sp} = [Pb^{2+}][SO_4^{2-}]$$

$$6.3 \times 10^{-7} = (0.0500 M)[SO_4^{2-}]$$

$$1.3 \times 10^{-5} = [SO_4^{2-}]$$

$$[Na_2SO_4] = [SO_4^{2-}] = 1.3 \times 10^{-5} M$$

$$\text{mol } Na_2SO_4 = (1.3 \times 10^{-5} M)(1.00 L) = 1.3 \times 10^{-5} \text{ mol}$$

$$\text{mass } Na_2SO_4 = (1.3 \times 10^{-5} \text{ mol})(142.1 \text{ g/mol})$$

$$= 1.8 \times 10^{-3} M$$

Assignment :

- Read sections 19-14 and 19-15
- Questions Page 554 #3-8