Totally Epic AP Chem Review: Solubility in a Day!

K_{sp}: Solubility product constant

- Equilibrium expression for the dissolution of a Solid.
- Like all K values, this is constant (at a constant temperature).
- Because solids are not included in an equilibrium expression, this will have only products.

Write the Ksp expression for each of the following dissolutions:

Salt	Dissociation reaction	K_{sp} Expression
K_2CO_3	$K_2CO_{3(s)} \rightarrow 2K^+_{(aq)} + CO_3^{2-}{}_{(aq)}$	$K_{sp} = [K^+]^2[CO_3^{2-}]$
Al_2S_3	$Al_2S_{3(s)} \rightarrow 2Al^{3+}_{(aq)} + 3S^{2-}_{(aq)}$	$K_{sp} = [Al^{3+}]^2[S^{2-}]^3$

Remember, there are some basic solubility rules you MUST know!

Always soluble: alkali metal cations, NH_4^+, and NO_3^-

Solubility "S" (aka Molar Solubility) = "x" in your K_{sp} RICE Table

How much of a Solid will dissolve per 1.0 L of solution (Units: M = mol/L)

Solubility is an equilibrium position and therefore can change (for example, if you change the number of ions in solution, this will shift the equilibrium position and thus, the solubility).

- Larger molar solubility values suggest more dissociation into ions ~~and greater solubility~~ more solid dissolves.
- Smaller molar solubility values suggest less dissociation into ions ~~and lower solubility~~ less solid dissolves.

1. The molar solubility of barium fluoride is at 25°C is $\underbrace{2.45 \times 10^{-5}}_{x}$. Calculate K_{sp}.

$$BaF_{2(s)} \rightleftharpoons Ba^{2+}_{(aq)} + 2F^-_{(aq)}$$

$+x \quad +2x$

$$K_{sp} = [Ba^{2+}][F^-]^2 = x(2x)^2 = 4x^3 = 4(2.45E-5)^3$$
$$= \boxed{5.88 \times 10^{-14}}$$

2. Calculate the molar solubility of nickel (II) carbonate, which has a K_{sp} of 1.4×10^{-7} at 25°C.

$$NiCO_{3(s)} \rightleftharpoons Ni^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$$

$+x \quad +x$

$$K_{sp} = [Ni^{2+}][CO_3^{2-}] = x^2 = 1.4E-7$$
$$x = \sqrt{1.4E-7} = \boxed{3.7 \times 10^{-4} M}$$

Precipitation: Will it or won't it?

Determining if Precipitation Will Occur: A Task for K vs Q!

Precipitation occurs when the concentrations of ions is greater than the solubility of the ionic compound.

$K_{sp} < Q$ more ions than the system can dissolve; precipitate will form!

$K_{sp} = Q$ **exactly** as many ions in solution as the system can dissolve; no precipitate

$K_{sp} > Q$ more of the solid will dissolve, and more ions will form; no precipitate forms

Important Ideas to Note:

1. If any solid is present, the solution is at equilibrium (a Saturated solution)
2. Ion concentration, [ions], is **independent** of volume when at equilibrium (for instance, in a Saturated solution).
3. If ions are present that could form multiple salts, the solid with the Smallest K_{sp} will form.

Let's Practice!

1. A chemist makes a 2.0 L saturated solution of $Ba_3(PO_4)_2$ solution, which has a $K_{sp} = 6.0 \times 10^{-39}$.
   a. What is the concentration of Ba^{2+} ions in solution?

$$K_{sp} = [Ba^{2+}]^3[PO_4^{3-}]^2 = (3x)^3(2x)^2 = 108x^5 = 6.0E-39$$

$$x = \sqrt[5]{\frac{6.0E-39}{108}} = 8.9E-9\,M \Rightarrow [Ba^{2+}] = 3x = 3(8.9E-9)$$
$$= \boxed{2.7 \times 10^{-8} M}$$

   b. After two days of sitting on the counter, some liquid has evaporated from the solution. Did $[Ba^{2+}]$ increase, decrease, or remain the same? Justify your answer.

$[Ba^{2+}]$ remains the Same! Although less liquid is present, the sol'n was already saturated, thus $[Ba^{2+}]$ can't increase.

   c. The chemist adds 3.00 g of solid $(NH_4)_3PO_4$ to the original saturated solution of $Ba_3(PO_4)_2$. Did $[Ba^{2+}]$ increase, decrease, or remain the same? Justify your answer.

$$Ba_3(PO_4)_{2(s)} \rightleftharpoons 3Ba^{2+}_{(aq)} + 2PO_4^{3-}{}_{(aq)}$$

$[Ba^{2+}]$ will decrease, b/c the added $(NH_4)_3PO_4$ will dissolve, $\uparrow [PO_4^{3-}]$, will will cause the rxn to shift left to re-establish equilibrium. This means more $Ba_3(PO_4)_{2\,(s)}$ will form, removing some $Ba^{2+}_{(aq)}$ from Sol'n.

The Common Ion Effect

Remember: solubility can change if you change reaction conditions!

- Le Châtelier's principle predicts that a salt will become less soluble in a solution that already contains one of its own ions already dissolved: what's known as a Common ion.
- The presence of a common ion acts like increasing the concentration of a product ion in the salt dissolution, causing the system to shift left to establish equilibrium (towards the Solid side).

Example:

in equimolar amounts

1. Circle any of the following compounds that contain a common ion to $MgCl_2$:

AgF_2 $NaCl$ $Mg(OH)_2$ $AlCl_3$ Al_2S_3

2. Which of the compounds above, if present in solution, would reduce the solubility of $MgCl_2$:
   a. the most? Why? $AlCl_3$, b/c it produces the most common ions per formula unit ($3\ Cl^-$)
   b. the least? Why? AgF_2 or Al_2S_3, b/c they contain no ions common to $MgCl_2$

The Effect of pH on Solubility

The common ion effect predicts that when a salt contains ions that can act as an acid or a base, the solubility of that salt will be affected by changes in pH.

Example: $Fe(OH)_3 \rightleftharpoons Fe^{3+}(aq) + 3\ OH^-(aq)$

1. Will iron (III) hydroxide be more, less, or equally soluble in a **basic** solution (when compared to its solubility in pure water)? Explain.

Less soluble, b/c OH^- ions are present in a basic sol'n, and OH^- is a common ion to $Fe(OH)_3$, thus decreasing the amount of $Fe(OH)_3$ which can dissolve.

2. Will iron (III) hydroxide be more, less, or equally soluble in an **acidic** solution (when compared to its solubility in pure water)? Explain.

More soluble, b/c the H^+ present in an acidic sol'n will react with the OH^- ions, removing them from sol'n. The decreased $[OH^-]$ will cause the rxn to shift right to re-establish equilibrium, allowing more $Fe(OH)_3$ to dissolve.