

## 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  $K_{sp}$  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  $NO_3^-$

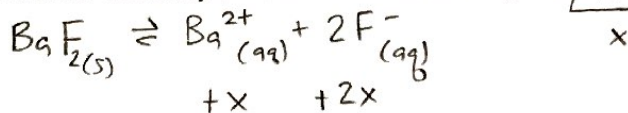
### Solubility "S" (aka Molar Solubility) = "x" in your $K_{sp}$ RICE Table

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*

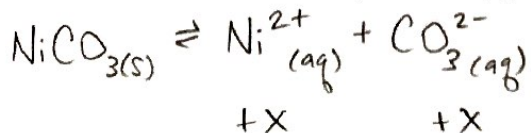
1. The molar solubility of barium fluoride is at 25°C is  $2.45 \times 10^{-5}$ . Calculate  $K_{sp}$ .



$$K_{sp} = [Ba^{2+}][F^-]^2 = x(2x)^2 = 4x^3 = 4(2.45E-5)^3$$

$$= \boxed{5.88 \times 10^{-14}}$$

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



$$K_{sp} = [Ni^{2+}][CO_3^{2-}] = x^2 = 1.4E-7$$

$$x = \sqrt{1.4E-7} = \boxed{3.7 \times 10^{-4} M}$$

## 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?

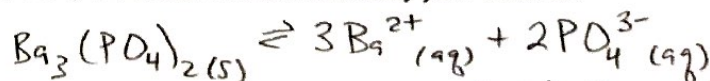
$$K_{sp} = [Ba^{2+}]^3 [PO_4^{3-}]^2 = (3x)^3 (2x)^2 = 108x^5 = 6.0E-39$$

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

- 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.



$[Ba^{2+}]$  will decrease, b/c the added  $(NH_4)_3PO_4$  will dissolve,  $\uparrow [PO_4^{3-}]$ , 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:

1. Circle any of the following compounds that contain a common ion to  $\text{MgCl}_2$ :

$\text{AgF}_2$

$\text{NaCl}$

$\text{Mg(OH)}_2$

$\text{AlCl}_3$

$\text{Al}_2\text{S}_3$

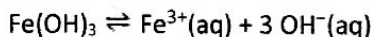
2. Which of the compounds above, if present in solution, would reduce the solubility of  $\text{MgCl}_2$ :

- a. the most? Why?  $\text{AlCl}_3$ , b/c it produces the most common ions per formula unit (3  $\text{Cl}^-$ )
- b. the least? Why?  $\text{AgF}_2$  or  $\text{Al}_2\text{S}_3$ , b/c they contain no ions common to  $\text{MgCl}_2$

## 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:



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  $\text{OH}^-$  ions are present in a basic sol'n, and  $\text{OH}^-$  is a common ion to  $\text{Fe(OH)}_3$ , thus decreasing the amount of  $\text{Fe(OH)}_3$  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  $\text{H}^+$  present in an acidic sol'n will react with the  $\text{OH}^-$  ions, removing them from sol'n. The decreased  $[\text{OH}^-]$  will cause the rxn to shift right to re-establish equilibrium, allowing more  $\text{Fe(OH)}_3$  to dissolve.