

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: \uparrow [product ions] = less salt will dissolve!

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

in equal amounts

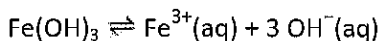
2. Which of the compounds above, if present in solution, would reduce the solubility of MgCl_2 :

- 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:



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

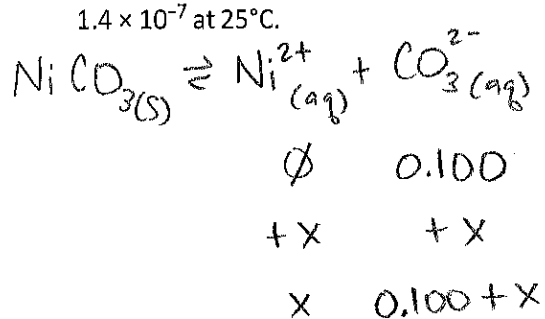
Less soluble. OH^- ions are present in a basic solution, which is a common ion to Fe(OH)_3 , so less Fe(OH)_3 will be able to dissolve in a basic solution.

4. 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! The H^+ present in an acidic sol'n will react w/ the OH^- ions (producing H_2O). This will reduce the concentration of OH^- ions, so the reaction will need to shift right to reach equilibrium, causing more Fe(OH)_3 to dissolve.

Common Ion Calculations

5. Calculate the molar solubility of nickel (II) carbonate in a solution containing 0.100 M NaCO_3 . The K_{sp} of NiCO_3 is 1.4×10^{-7} at 25°C.



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

$$= x(0.100 + x) \approx 0.100x = 1.4 \times 10^{-7}$$

$0.100 > 1000(K_{sp}) \Rightarrow x \text{ negligible!}$
 \uparrow
 1.4×10^{-7}

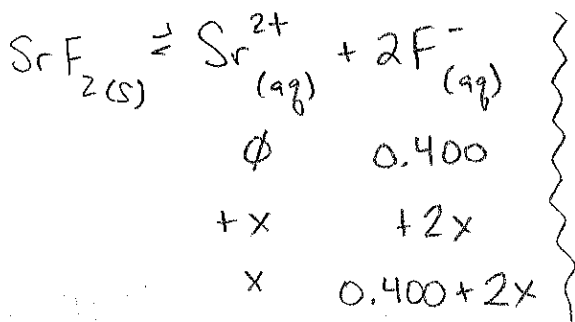
$$x = \frac{1.4 \times 10^{-7}}{0.100} = \boxed{1.4 \times 10^{-6} \text{ M}}$$

molar solubility in H_2O ?

$$K_{sp} = x^2$$

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

6. Calculate the molar solubility of SrF_2 in a solution containing 0.400 M NaF . The K_{sp} of SrF_2 is 7.9×10^{-10} at 25°C.



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

$$= x(0.400 + 2x)^2 \approx x(0.400)^2 = 0.160x$$

$0.400 > 1000(K_{sp}) \Rightarrow x \text{ negligible}$
 \uparrow
 7.9×10^{-10}

$$7.9 \times 10^{-10} = 0.160x \Rightarrow x = \frac{7.9 \times 10^{-10}}{0.160} = \boxed{4.9 \times 10^{-9} \text{ M}}$$

----- Common Ion Applications: Achieving a Desired Concentration -----

You can use the common ion effect to control the concentration of the non-common ion in solution! Let's see how this chemical magic works. **★ no RICE needed!**

7. In pure water at 25°C, the K_{sp} of CaCO_3 is 3.8×10^{-9} . $\text{Ca}(\text{NO}_3)_2$ is added to 1.00 L of a saturated solution of CaCO_3 at 25°C until the $[\text{CO}_3^{2-}]$ is reduced to 2.3×10^{-7} M. How many moles of $\text{Ca}(\text{NO}_3)_2$ are dissolved in solution at the point when $[\text{CO}_3^{2-}] = 2.3 \times 10^{-7}$ M? (Assume the added $\text{Ca}(\text{NO}_3)_2$ has a negligible effect on the total volume of solution.)

Step 1:
find $[\text{Ca}^{2+}]$

$$K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}] = [\text{Ca}^{2+}](2.3 \times 10^{-7}) = 3.8 \times 10^{-9}$$

$$\Rightarrow [\text{Ca}^{2+}] = \frac{3.8 \times 10^{-9}}{2.3 \times 10^{-7}} = 1.7 \times 10^{-2} \text{ M Ca}^{2+}$$

Step 2: find moles!

$$1.7 \times 10^{-2} \text{ M Ca}^{2+} \times 1.00 \text{ L} = 1.7 \times 10^{-2} \text{ mol Ca}^{2+} \times \frac{1 \text{ mol Ca}(\text{NO}_3)_2}{1 \text{ mol Ca}^{2+}}$$

$$= \boxed{1.7 \times 10^{-2} \text{ mol Ca}(\text{NO}_3)_2}$$

8. Given a 2.00 L saturated solution of $\text{Cu}(\text{IO}_3)_2$, $K_{sp} = 1.4 \times 10^{-7}$, how many moles of NaIO_3 would need to be dissolved in solution to reduce $[\text{Cu}^{2+}]$ to $6.0 \times 10^{-5} \text{ M}$? (Assume the added NaIO_3 does not appreciably change the total volume of solution.)

$$K_{sp} = [\text{Cu}^{2+}][\text{IO}_3^-]^2 = (6.0 \times 10^{-5})[\text{IO}_3^-]^2 = 1.4 \times 10^{-7}$$

$$[\text{IO}_3^-] = \sqrt{\frac{1.4 \times 10^{-7}}{6.0 \times 10^{-5}}} = 0.048 \text{ M } \text{IO}_3^- \times 2.00 \text{ L} = 0.097 \text{ mol } \text{IO}_3^-$$

$$0.097 \text{ mol } \text{IO}_3^- \times \frac{1 \text{ mol } \text{NaIO}_3}{1 \text{ mol } \text{IO}_3^-} = \boxed{0.097 \text{ mol } \text{NaIO}_3}$$

Even more practice!

9. Copper(I) bromide has a measured solubility of $2.0 \times 10^{-4} \text{ mol/L}$ at 25°C . Calculate its K_{sp} value.

$$K_{sp} = [\text{Cu}^+][\text{Br}^-] = x^2 = (2.0 \times 10^{-4})^2 = \boxed{4.0 \times 10^{-8}}$$

10. The K_{sp} value for copper(II) iodate, $\text{Cu}(\text{IO}_3)_2$, is 1.4×10^{-7} at 25°C . What is the maximum mass, in grams, of copper (II) iodate that can dissolve in 500. mL of water?

$$K_{sp} = [\text{Cu}^{2+}][\text{IO}_3^-]^2 = x(2x)^2 = 4x^3 = 1.4 \times 10^{-7}$$

$$x = \sqrt[3]{\frac{1.4 \times 10^{-7}}{4}} = \underbrace{3.271 \times 10^{-3} \text{ M}}_x \times 0.500 \text{ L} = 1.636 \times 10^{-3} \text{ mol } \text{Cu}(\text{IO}_3)_2 \times \frac{413.37 \text{ g}}{1 \text{ mol}} = \boxed{0.68 \text{ g } \text{Cu}(\text{IO}_3)_2}$$

11. In pure water at 25°C , the molar solubility of PbCl_2 is 1.3×10^{-6} and the K_{sp} is 1.6×10^{-5} . LiCl is added to 5.00 L of a saturated solution of PbCl_2 at 25°C until the $[\text{Pb}^{2+}]$ is reduced to $4.5 \times 10^{-4} \text{ M}$. How many moles of chloride ions are dissolved in solution at the point when $[\text{Pb}^{2+}] = 4.5 \times 10^{-4} \text{ M}$? (Assume the added $\text{Ca}(\text{NO}_3)_2$ has a negligible effect on the total volume of solution.) LiCl

$$K_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]^2 = (4.5 \times 10^{-4})[\text{Cl}^-]^2 = 1.6 \times 10^{-5}$$

$$[\text{Cl}^-] = \sqrt{\frac{1.6 \times 10^{-5}}{4.5 \times 10^{-4}}} = 0.19 \text{ M } \text{Cl}^- \times 5.00 \text{ L} = \boxed{0.94 \text{ mol } \text{Cl}^-}$$