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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 know 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: ↑[product ions] = less salt will dissolve!

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

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

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.

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

	Ni^{2+}	CO_3^{2-}
	∅	0.100
	+x	+x
	x	0.100 + x

$$K_{sp} = [Ni^{2+}][CO_3^{2-}]$$

$$= x\underbrace{(0.100 + x)} \approx 0.100x = 1.4E-7$$

$0.100 > 1000(K_{sp}) \Rightarrow x$ negligible!

↑ 1.4E-7

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

molar solubility in H_2O?

$K_{sp} = x^2$

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

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.

$$SrF_{2(s)} \rightleftharpoons Sr^{2+}_{(aq)} + 2F^-_{(aq)}$$

	Sr^{2+}	F^-
	∅	0.400
	+x	+2x
	x	0.400 + 2x

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

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

$0.400 > 1000(K_{sp}) \Rightarrow x$ negligible

↑ 7.9E-10

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

Common Ion Applications: Achieving a Desired Concentration

You can use the common ion effect to control the <u>concentration</u> 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}. $Ca(NO_3)_2$ is added to 1.00 L of a saturated solution of $CaCO_3$ at 25°C until the $[CO_3^{2-}]$ is reduced to 2.3×10^{-7} M. How many moles of $Ca(NO_3)_2$ are dissolved in solution at the point when $[CO_3^{2-}] = 2.3 \times 10^{-7}$ M? (Assume the added $Ca(NO_3)_2$ has a negligible effect on the total volume of solution.)

Step 1: find $[CO_3^{2-}]$

$$K_{sp} = [Ca^{2+}][CO_3^{2-}] = [Ca^{2+}](2.3E-7) = 3.8E-9$$

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

Step 2: find moles!

$$1.7E-2\ M\ Ca^{2+} \times 1.00\ L = 1.7E-2\ mol\ Ca^{2+} \times \frac{1\ mol\ Ca(NO_3)_2}{1\ mol\ Ca^{2+}}$$

$$= \boxed{1.7 \times 10^{-2}\ mol\ Ca(NO_3)_2}$$

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8. Given a 2.00 L saturated solution of $Cu(IO_3)_2$, $K_{sp} = 1.4 \times 10^{-7}$, how many moles of $NaIO_3$ would need to be dissolved in solution to reduce $[Cu^{2+}]$ to 6.0×10^{-5} M? (Assume the added $NaIO_3$ does not appreciably change the total volume of solution.)

$$K_{sp} = [Cu^{2+}][IO_3^-]^2 = (6.0E-5)[IO_3^-]^2 = 1.4E-7$$

$$[IO_3^-] = \sqrt{\frac{1.4E-7}{6.0E-5}} = 0.048\ M\ IO_3^- \times 2.00\ L = 0.097\ mol\ IO_3^-$$

$$0.097\ mol\ IO_3^- \times \frac{1\ mol\ NaIO_3}{1\ mol\ IO_3^-} = \boxed{0.097\ mol\ NaIO_3}$$

Even more practice!

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

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

10. The K_{sp} value for copper(II) iodate, $Cu(IO_3)_2$, 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?

$$K_{sp} = [Cu^{2+}][IO_3^-]^2 = x(2x)^2 = 4x^3 = 1.4E-7$$

$$x = \sqrt[3]{\frac{1.4E-7}{4}} = \underbrace{3.271E-3\ M}_{x} \times 0.500\ L = 1.636E-3\ mol\ Cu(IO_3)_2 \times \frac{413.37\ g}{1\ mol}$$

$$= \boxed{0.68\ g\ Cu(IO_3)_2}$$

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 $[Pb^{2+}]$ is reduced to 4.5×10^{-4} M. How many moles of chloride ions are dissolved in solution at the point when $[Pb^{2+}] = 4.5 \times 10^{-4}$ M? (Assume the added ~~$Ca(NO_3)_2$~~ LiCl has a negligible effect on the total volume of solution.)

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

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