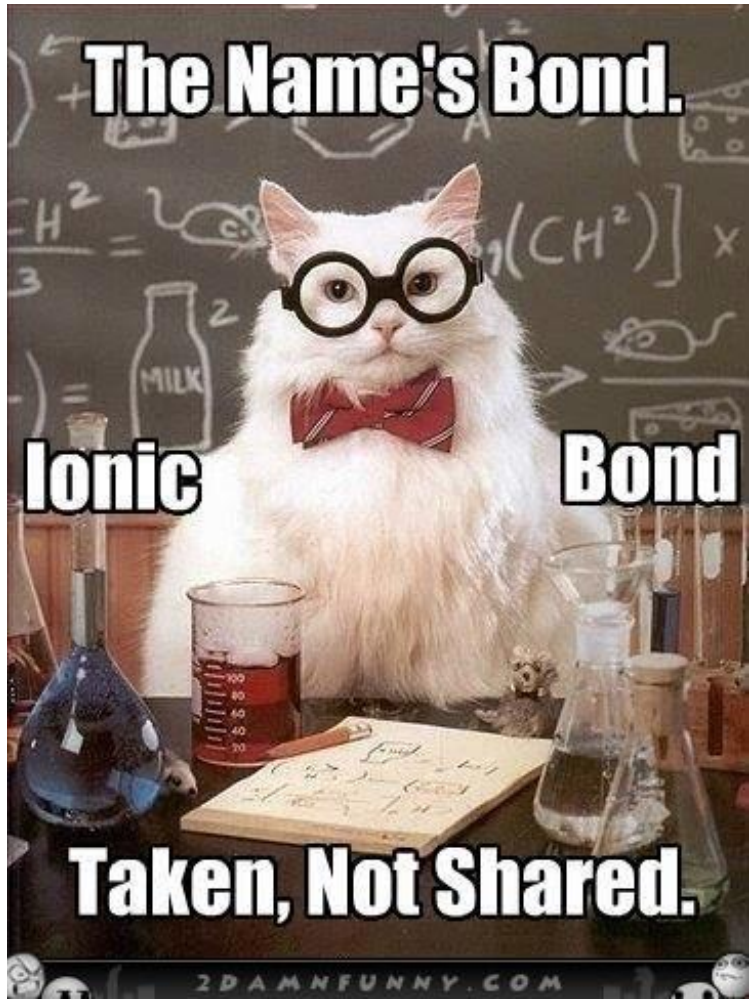
Common Ion Effect



Outcome:

Solve problems involving Ksp

Common Ion Effect…

Common Ions:

- When an **IONIC** compound dissolves in **PURE** water, the **INITIAL CONCENTRATION** of the **IONS** is **ZERO**.
- If an ionic compound were to dissolve in a solution that **ALREADY CONTAINS** an **ION** that is **COMMON** to the **SOLUTE**, this is **NOT** the **CASE**.
- Even though the **INITIAL** concentrations may **NOT** be **ZERO**, the **PRODUCT** of the **IONS** must still **EQUAL** the **Ksp** at **EQUILIBRIUM**.

Example:

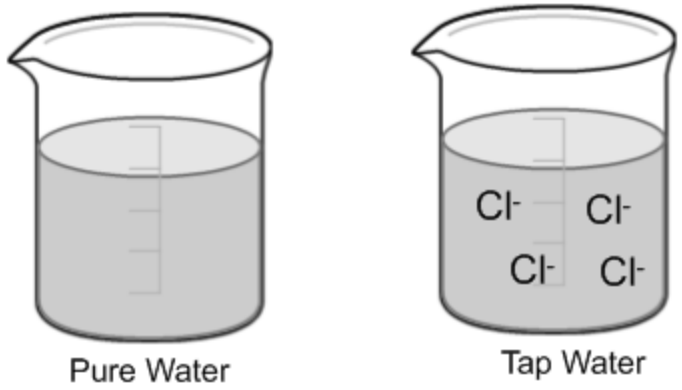
How would the solubility of AgCl in pure water change if we were to dissolve it in tap water?

Let's write the dissociation equation:

$$AgCl_{(s)} \leftrightarrows Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$$

Common Ion Effect…

Tap water has chlorine (Cl^-) added to kill bacteria, so when we dissolve AgCl in water, there are already some Cl^- ions present.



$$AgCl_{(s)} \leftrightarrows Ag^+_{(aq)} + Cl^-_{(aq)}$$

According to Le Chatelier, as we add Cl- ions the equilibrium would shift to the left, to use up the excess product.

Therefore:

→ ***More solid would be formed, and solubility would decrease.***

Common Ion Examples…

1. Determine the solubility of AgCl in pure water, and in a solution of 0.10 mol/L NaCl. (Ksp of AgCl = 1.7×10^{-10})

Pure H_2O

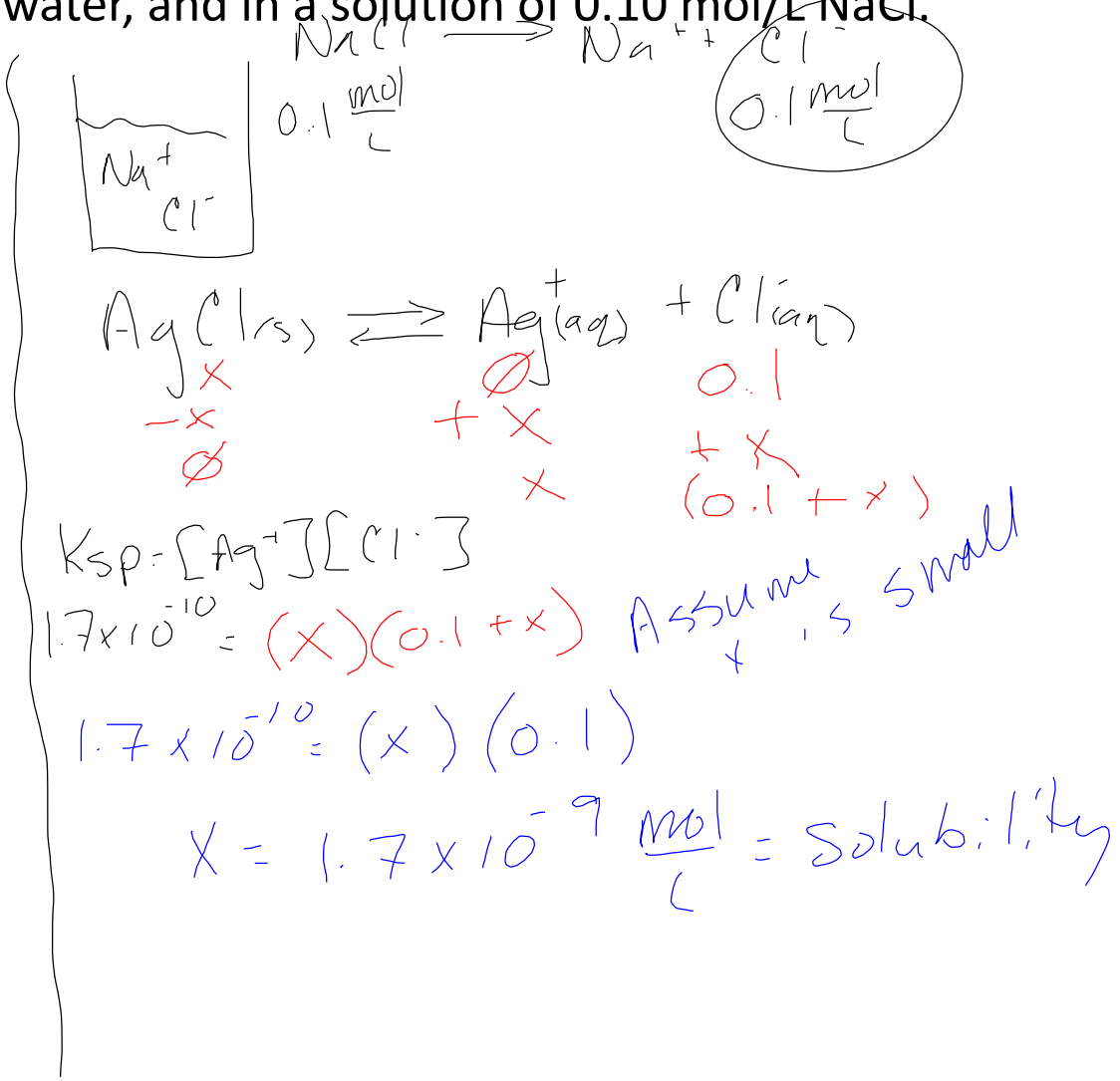
$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$

x	0	0
−x	+x	+x
0	x	x

$K_{sp} = [Ag^+][Cl^-]$

$1.7 \times 10^{-10} = (x)(x)$

$x = 1.3 \times 10^{-5} \frac{mol}{L}$ = solubility



Common Ion Examples…

2. The Ksp of lead (II) chloride is 1.6×10^{-5}. What is the solubility of lead (II) chloride in 0.10 mol/L magnesium chloride?

$4 \times 10^{-4} \frac{mol}{L}$

$6.32 \times 10^{-3} \frac{mol}{L}$

Try these ones…

1. The Ksp of lead (II) chloride is 1.6×10^{-5}. What is the solubility of lead (II) chloride in 0.10 mol/L lead(II) nitrate?

Try these ones…

1. The Ksp of calcium hydroxide is 5.5×10^{-6}. Calculate the solubility of calcium hydroxide when it is dissolved in:

   a) Pure water:

$Ca(OH)_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2OH^-_{(aq)}$

	Ca^{2+}	OH^-
X	∅	∅
−x	+x	+2x
∅	x	2x

$K_{sp} = [Ca^{2+}][OH^-]^2$

$5.5 \times 10^{-6} = (x)(2x)^2$

$5.5 \times 10^{-6} = 4x^3$

$\sqrt[3]{\frac{5.5 \times 10^{-6}}{4}} = \sqrt[3]{x^3}$

$x = 0.011 \frac{mol}{L}$ = solubility

   b) 0.2M $Mg(OH)_2$ → Mg^{2+} + $2OH^-$

Mg^{2+}: 0.2 $\frac{mol}{L}$

$2OH^-$: 0.4 $\frac{mol}{L}$

$Ca(OH)_{2(s)} \rightleftharpoons Ca^{2+} + 2OH^-$

	Ca^{2+}	OH^-
X	∅	0.4
−x	+x	+2x
∅	(x)	(0.4+2x)

$K_{sp} = [Ca^{2+}][OH^-]^2$

$5.5 \times 10^{-6} = (x)(0.4+2x)^2$ assume x is small

$5.5 \times 10^{-6} = (x)(0.4)^2$

$x = 3.4 \times 10^{-5} \frac{mol}{L}$ = solubility