$$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$$

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

↪ This equilibrium constant is given a special name - the SOLUBILITY PRODUCT CONSTANT - because the equilibrium expression for the dissolving of a salt always appears as a PRODUCT of the concentrations of the ions in the compound!

Remember, Ksp is an equilibrium constant, so everything that applies to equilibrium constants applies to the solubility constant - including what to do with coefficients:

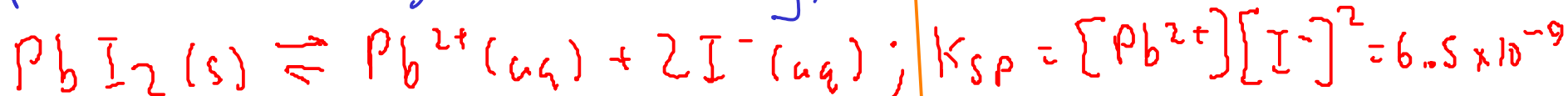What is the solubility product constant expression for calcium phosphate?

Ca^{2+} PO_4^{3-}

$$Ca_3(PO_4)_2 (s) \rightleftharpoons 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)$$
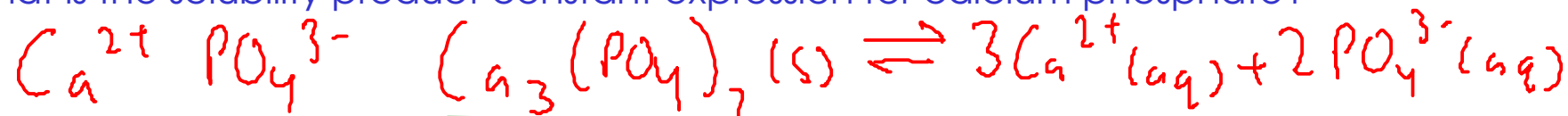
$$K_{SP} = [Ca^{2+}]^3[PO_4^{3-}]^2$$

Solubility calculations and Ksp

You can calculate the solubility of a compound if you know Ksp!

Calculate the solubility (in g/L) of lead(II) iodide at 25C.

$$K_{sp} = 6.5 \times 10^{-9} \quad ; \quad FW = 461.0 \, g/mol$$

$$PbI_2 (s) \rightleftharpoons Pb^{2+}(aq) + 2I^-(aq) ; \quad \boxed{K_{sp} = [Pb^{2+}][I^-]^2 = 6.5 \times 10^{-9}}$$

We need to solve this equilibrium expression to find out how much of the compound has dissolved!

Species	[Initial]	Δ	[Equil]
Pb^{2+}	0	$+x$	x
I^-	0	$+2x$	$2x$

$$(x)(2x)^2 = 6.5 \times 10^{-9}$$
$$4x^3 = 6.5 \times 10^{-9}$$

$$x = 0.0011756673$$
$$x = [Pb^{2+}] = [PbI_2] \, dissolved$$

Since the dissolved lead concentration equals the dissolved lead iodide concentration...

$$\frac{0.0011756673 \, mol \, PbI_2}{L} \times \frac{461.0 g \, PbI_2}{mol \, PbI_2} = \boxed{0.54 \, g/L} = 540 \, ppm \, PbI_2 \left(\frac{mg}{L}\right)$$

If you wanted to know ppm lead ...

$$\frac{0.0011756673 \, mol \, PbI_2}{L} \times \frac{207.2 g \, Pb}{mol \, Pb} = 0.24 \, g/L = 240 \, ppm \, Pb$$

Precipitation - also known as the reaction quotient

To predict whether a salt at a given concentration will precipitate out, calculate the reaction quotient Q and compare it to the Ksp

Example: $$AgCl \ (s) \rightleftharpoons Ag^+ (aq) + Cl^- (aq) ; \ K_{sp} = 1.8 \times 10^{-10}$$
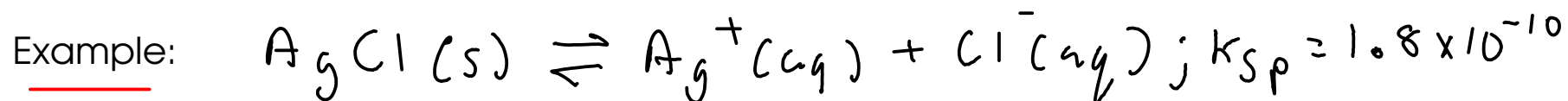
$$Q = [Ag^+][Cl^-]$$

IF...

* Q < Ksp ; the reaction proceeds to produce more products (dissolved ions), so more solid is able to dissolve: NO PRECIPITATION

* Q > Ksp ; the reaction proceeds to produce more reactants (solid), so solid falls out of solution: PRECIPITATION OCCURS

* Q = Ksp ; the reaction is at equilibrium. PRECIPITATION IS JUST BEGINNING

Would a solution with (Ag+) = 0.014 M and (Cl-) = 0.00042 M precipitate?

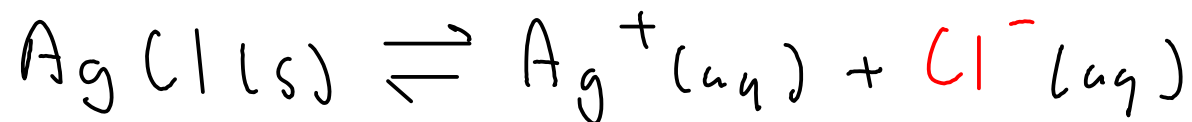$$Q = [Ag^+][Cl^-] = (0.014)(0.00042) = 5.88 \times 10^{-6}$$

$$5.88 \times 10^{-6} > 1.8 \times 10^{-10}$$

$$Q \qquad > \qquad Ksp$$

Since Q > Ksp, PRECIPITATION OCCURS!

Le Chateleir's Principle

The "common ion effect" affects the solubility of a compound in solution. The presence of one of the ions in a salt in the solution will REDUCE THE SOLUBILITY of that salt!
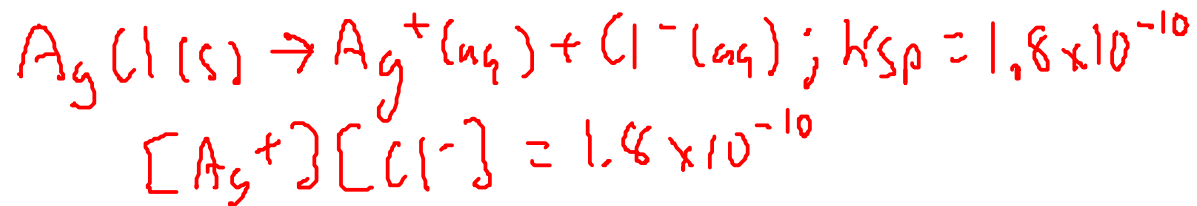
$$AgCl (s) \rightleftharpoons Ag^+ (aq) + Cl^- (aq)$$

Silver chloride is much less soluble in a solution of 0.1 M NaCl than it is in distilled water. Why? The presence of CHLORIDE ION forces the solubility equilibrium back to the left, meaning less silver chloride can dissolve!

Solubility can also be affected by pH - depending on the acidic or basic properties of the salt!

A second example: Salicylic acid in the characterization lab

Calculate the solubility of AgCl in distilled water. Then, calculate the solubility of AgCl in 0.10 M NaCl solution. Report both answers in parts per million (mg/L)

$$AgCl_{(s)} \rightarrow Ag^+_{(aq)} + Cl^-_{(aq)} \; ; \; K_{sp} = 1.8 \times 10^{-10}$$

$Ag: 1 \times 107.9$

$Cl: 1 \times 35.45$

$$[Ag^+][Cl^-] = 1.8 \times 10^{-10}$$

143.35 g/mol

For distilled water:

Species	[Init]	Δ	[Equil]
Ag^+	O	$+X$	X
Cl^-	O	$+X$	X

$$x^2 = 1.8 \times 10^{-10}$$
$$x = 1.341640786 \times 10^{-5} M = [Ag^+] = [AgCl]_{dissolved}$$

$$\frac{1.341640786 \times 10^{-5} \text{ mol } AgCl}{L} \times \frac{143.35 \text{ g } AgCl}{\text{mol } AgCl} = 0.0019 \text{ g}/L = \boxed{1.9 \text{ ppm}}$$

AgCl, in distilled water

$$AgCl(s) \rightarrow Ag^+(aq) + Cl^-(aq) \; ; \; K_{sp} = 1.8 \times 10^{-10}$$

$$[Ag^+][Cl^-] = 1.8 \times 10^{-10}$$
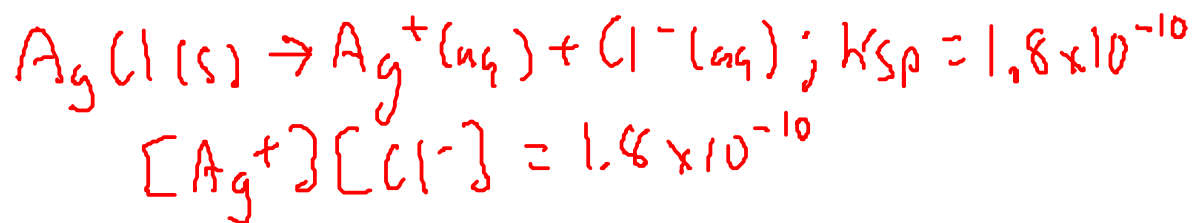
For solubility in 0.10 M NaCl:

Species	[Initial]	Δ	[Equil]
Ag^+	O	$+x$	x
Cl^-	0.10	$+x$	$(0.10 + x)$

$$(x)(0.10 + x) = 1.8 \times 10^{-10}$$

We know that x << 0.10, so ...

$$0.10 \, x = 1.8 \times 10^{-10}$$

$$x = 1.8 \times 10^{-9} = [Ag^+] = [AgCl]_{dissolved}$$

$$\frac{1.8 \times 10^{-9} \, mol \; AgCl}{L} \times \frac{143.35 \, g \, AgCl}{mol \; AgCl} = 2.6 \times 10^{-7} \, g/_L = \boxed{\begin{array}{c} 2.6 \times 10^{-4} \, ppm \\ (0.00026 \, ppm) \end{array}}$$
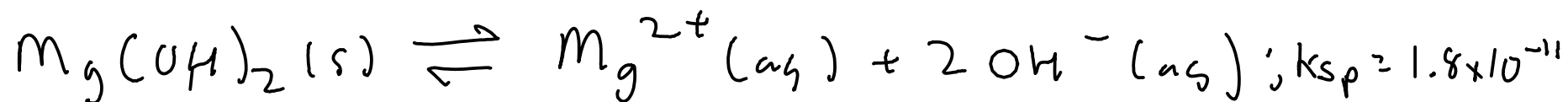
Compare:
1.9 ppm in distilled water
0.00026 ppm in 0.1M NaCl

... for AgCl in 0.10M NaCl solution

Conclusion: The presence of a common ion greatly decreases solubility.

pH AND SOLUBILITY

$$Mg(OH)_2 (s) \rightleftharpoons Mg^{2+} (aq) + 2 OH^- (aq) ; k_{sp} = 1.8 \times 10^{-11}$$

This compound's solubility is pH dependent. How?

> * In a BASIC solution, the concentration of hydroxide ion in solution is high.
> , so solubility is LOWER than in pure water.

> * In an ACIDIC solution, we have a significant amount of hydronium, which
> can react with hydroxide. This lowers the hydroxide concentration and makes
> magnesium hydroxide MORE SOLUBLE

Generalizing

If a compound is BASIC, then it will be LESS SOLUBLE in basic solutions, and MORE SOLUBLE in acidic solutions!

If a compound is ACIDIC, then it will be MORE SOLUBLE in basic solutions, and LESS SOLUBLE in acidic solutions!

If a compound is NEUTRAL (neither acidic nor basic), then its solubility will be UNAFFECTED by pH