

SOLUTION: Homogeneous mixture of substances Solutions contain:

SOLUTE: Component(s) of a solution present in small amount

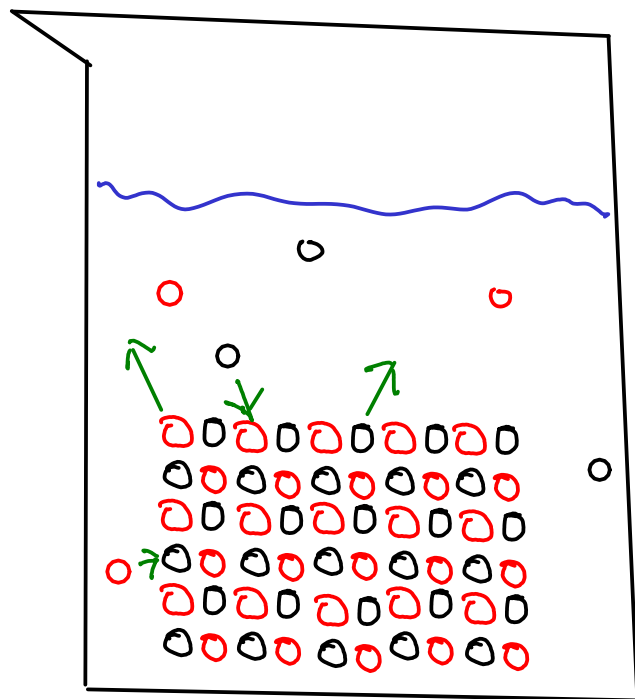
SOLVENT: Component of a solution present in greatest amount

We usually call water the solvent in aqueous mixtures, even if the water is present in smaller amount than another component

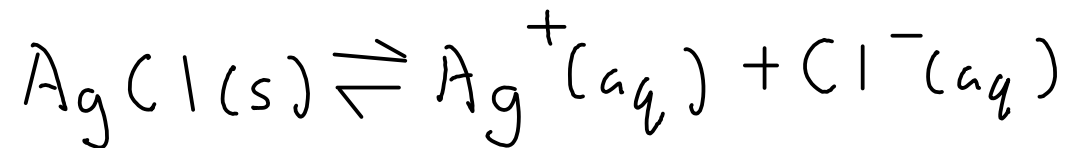
SOLUBILITY: The amount of a solute that will dissolve in a given volume of solvent

SATURATED SOLUTION: Contains the maximum amount of solute that it is possible to dissolve in a given volume of solvent!

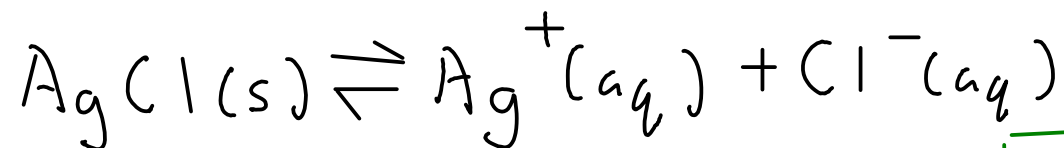
A SATURATED SOLUTION is a solution where dissolved solute exists in an EQUILIBRIUM with undissolved solute!



Example: Consider a saturated solution of silver chloride:

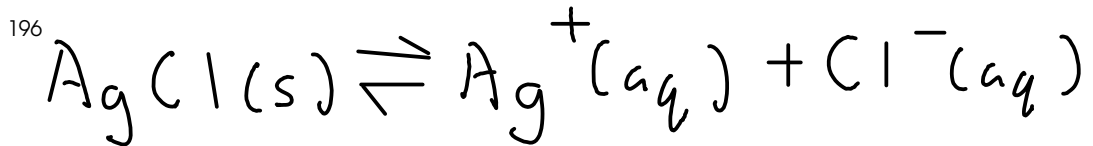


At equilibrium, the rate of dissolving equals the rate of crystallization!



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

... what does this equilibrium constant tell us? That silver chloride isn't very soluble!

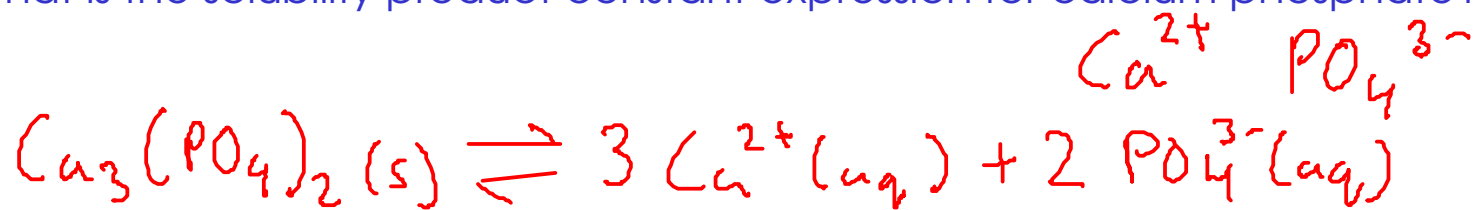


$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

↪ 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, K_{sp} 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?



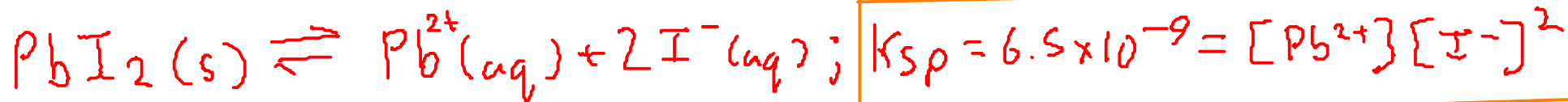
$$K_{sp} = [\text{Ca}^{2+}]^3 [\text{PO}_4^{3-}]^2$$

197 Solubility calculations and K_{sp}

You can calculate the solubility of a compound if you know K_{sp}!

Calculate the solubility (in g/L) of lead(II) iodide at 25°C. (see p A-15 in book)

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



We need to solve this equation, since these will give us equilibrium concentrations.

Species	[Initial]	Δ	[Equilibrium]
Pb ²⁺	0	+x	x
I ⁻	0	+2x	2x

$$6.5 \times 10^{-9} = (x)(2x)^2 \quad | \quad x = 0.0011756673 = [\text{Pb}^{2+}] = [\text{PbI}_2]_{\text{dissolved}}$$

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

Since the dissolved lead and dissolved lead(II) iodide concentrations are the same...

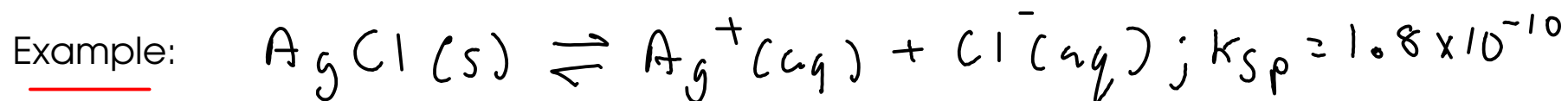
$$\frac{0.0011756673 \text{ mol PbI}_2}{\text{L}} \times \frac{461.0 \text{ g PbI}_2}{\text{mol PbI}_2} = 0.54 \text{ g/L} = 540 \text{ ppm PbI}_2 (\text{mg/L})$$

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

ppm (parts per million) is a common unit used for small concentrations. In dilute aqueous solutions, it's equal to mg/L

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 K_{sp}



$$Q = [Ag^+][Cl^-]$$

IF...

- * $Q < K_{sp}$; the reaction proceeds to produce more products (dissolved ions), so more solid is able to dissolve: NO PRECIPITATION
- * $Q > K_{sp}$; the reaction proceeds to produce more reactants (solid), so solid falls out of solution: PRECIPITATION OCCURS
- * $Q = K_{sp}$; the reaction is at equilibrium. PRECIPITATION IS JUST BEGINNING

Would a solution with $[Ag^+] = 0.014$ M and $[Cl^-] = 0.00042$ M precipitate?

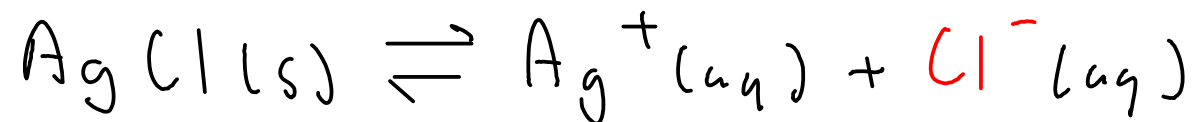
$$Q = [Ag^+][Cl^-] = (0.014)(0.00042) = 5.88 \times 10^{-6}$$

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

Q > K_{sp}, so precipitation occurs!

Le Chateleur'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!

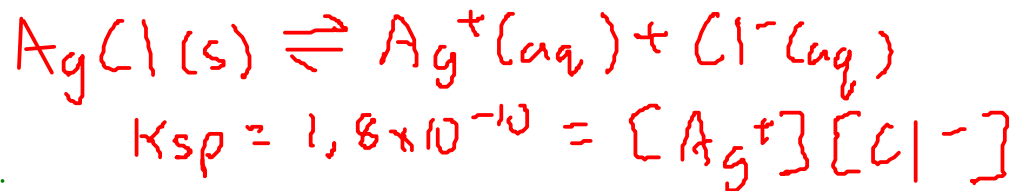


Silver chloride is much less soluble in a solution of 0.1 M NaCl than it is in distilled water. Why? The presence of CHLORIDE ION forces the solubility equilibrium back to the left, meaning less silver chloride can dissolve!

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)



For distilled water...

Species	[Initial]	Δ	[Equilibrium]
Ag^+	0	+x	x
Cl^-	0	+x	x

$$x^2 = 1.8 \times 10^{-10}$$

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

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

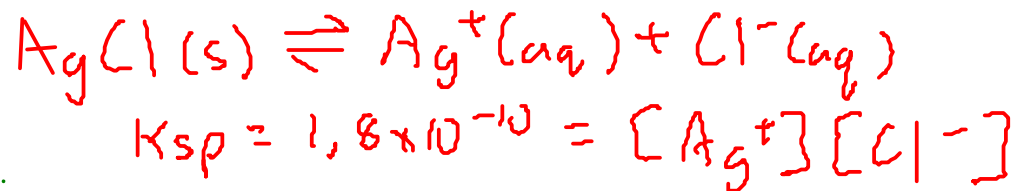
in distilled water

AgCl:

$$\text{Ag} : 1 \times 107.9$$

$$\text{Cl} : 1 \times 35.45$$

$$\hline 143.35 \text{ g/mol}$$



For 0.10 M NaCl solution...



Species	[Initial]	Δ	[Equilibrium]
Ag^+	0	+x	x
Cl^-	0.10	+x	0.10 + x

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

\downarrow $x \ll 0.10$, so $0.10 + x \approx 0.10$

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

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

$$\frac{1.8 \times 10^{-9} \text{ mol AgCl}}{1} \times \frac{143.35 \text{ g AgCl}}{\text{mol AgCl}} = 2.6 \times 10^{-7} \text{ g/L} = 2.6 \times 10^{-4} \text{ ppm}$$

(0.00026 ppm)

in 0.10 M NaCl

Compare:

1.9 ppm AgCl in distilled water

0.00026 ppm AgCl in aqueous 0.10 M NaCl solution

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