

SOLUTION: Homogeneous mixture of substances Solutions contain:

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

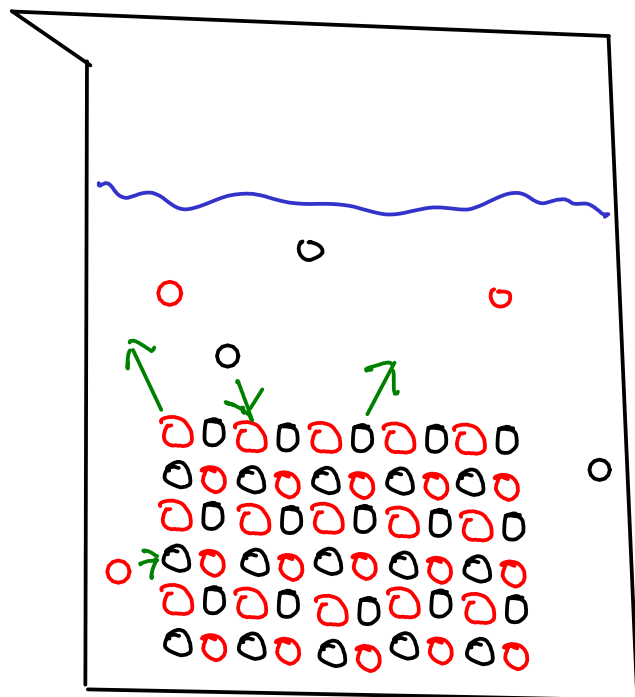
SOLVENT: Component of a solution present in greatest amount

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

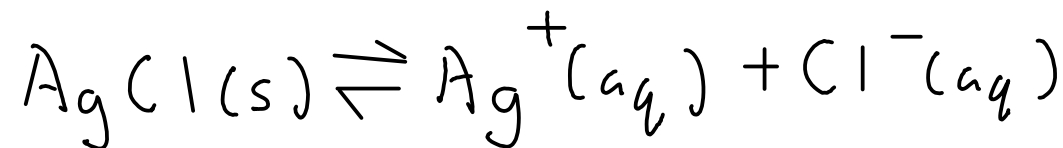
SOLUBILITY: The amount of a solute that will dissolve in a given volume of solvent

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

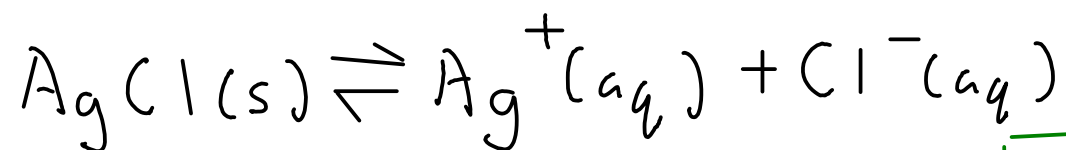
A SATURATED SOLUTION is a solution where dissolved solute exists in an EQUILIBRIUM with undissolved solute!



Example: Consider a saturated solution of silver chloride:

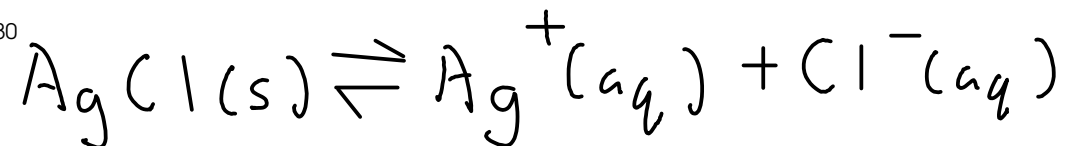


At equilibrium, the rate of dissolving equals the rate of crystallization!



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

... 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?



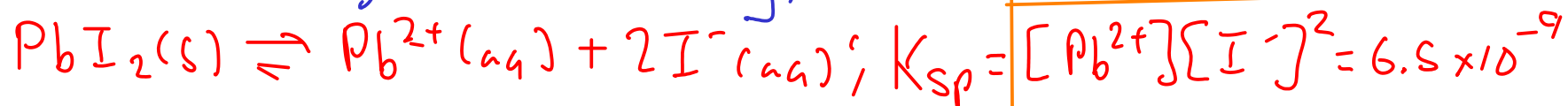
$$K_{sp} = [\text{Ca}^{2+}]^3[\text{PO}_4^{3-}]^2$$

181 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)

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



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

Solve to find equilibrium concentration of the ions ...

Define "x" as the change in lead(II) concentration...

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

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

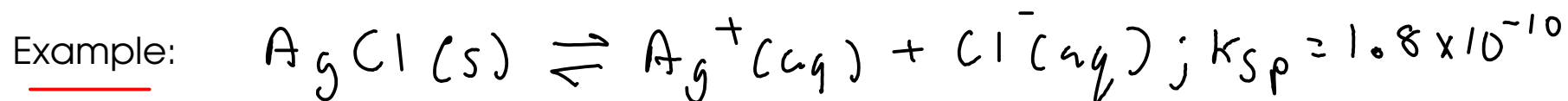
Since the dissolved lead(II) iodide concentration equals the lead(II) ion concentration, all we have to do is take the concentration and put it into the right units.

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

ppm: Parts Per Million. This is a common unit of solubility for "insoluble" or trace compounds in water. Same thing as mg/L for dilute aqueous solutions!

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 $(Ag^+) = 0.014 \text{ M}$ and $(Cl^-) = 0.00042 \text{ M}$ precipitate?

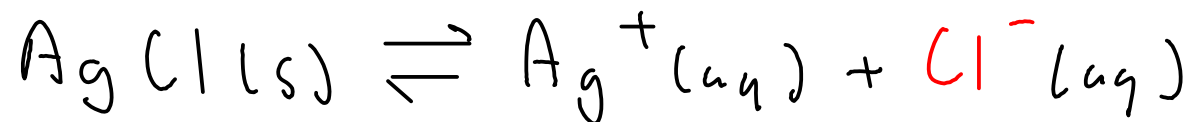
$$Q = [Ag^+][Cl^-] = (0.014)(0.00042) = 5.88 \times 10^{-6}$$

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

Since $Q > K_{sp}$, 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!



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 (FW = 143.35 g/mol) 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

Let "x" equal the change in dissolved Ag⁺ concentration.

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

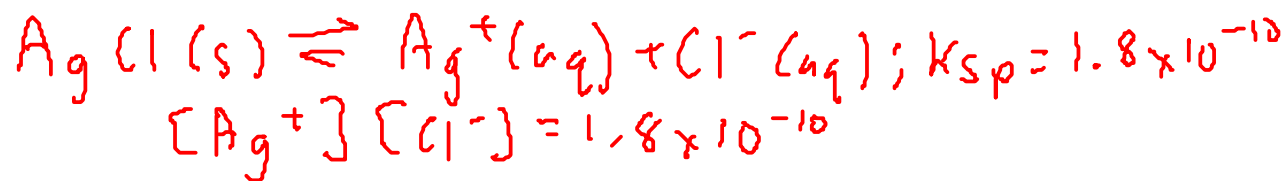
$$x^2 = 1.8 \times 10^{-10}$$

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

Convert answer's units to ppm (mg/L)...

$$\frac{1.341640786 \times 10^{-5} \text{ mol AgCl}}{\text{L}} \times \frac{143.35 \text{ g AgCl}}{\text{mol AgCl}} \times \frac{\text{mg}}{10^{-3} \text{ g}} = 1.9 \text{ ppm AgCl in distilled H}_2\text{O}$$

↑
converts mol to g
↑
converts g to mg



For solubility in 0.10 M NaCl solution...

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

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

↓ We know that $x \ll 0.10$, so $0.10 + x \approx 0.10$

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

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

$$\frac{1.8 \times 10^{-9} \text{ mol AgCl}}{1} \times \frac{143.35 \text{ g AgCl}}{\text{mol AgCl}} \times \frac{\text{mg}}{10^{-3} \text{ g}} = \boxed{2.6 \times 10^{-4} \text{ ppm AgCl in 0.10 M NaCl}}$$

(0.00026 ppm)

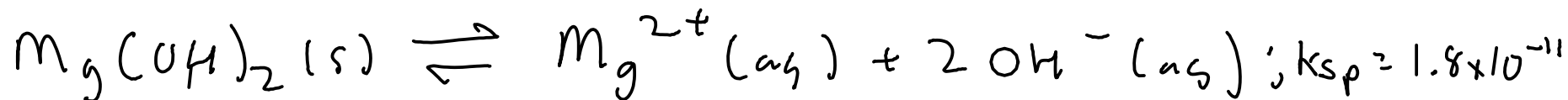
Compare:

1.9 ppm AgCl in distilled water vs

0.00026 ppm AgCl in 0.10 M NaCl

Conclusion: The presence of a common ion greatly decreases solubility

pH AND SOLUBILITY



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