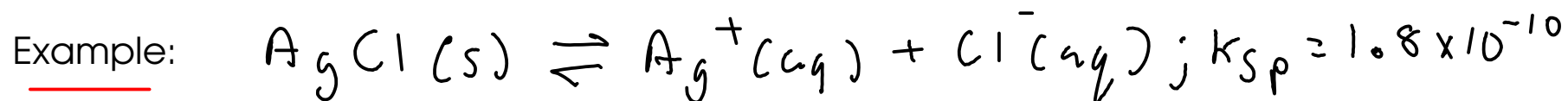


## Precipitation - also known as the reaction quotient

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?

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

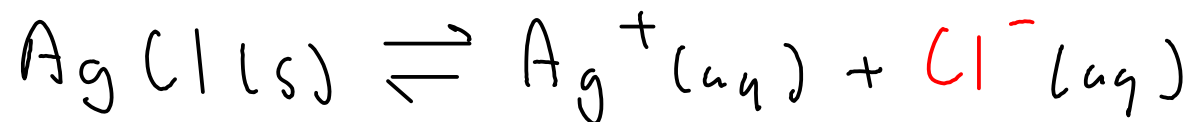
$$Q = 5.88 \times 10^{-6} > 1.8 \times 10^{-10}$$

$$Q > K_{sp}$$

Since  $Q > K_{sp}$ , precipitation occurs!

## <sup>191</sup> 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!

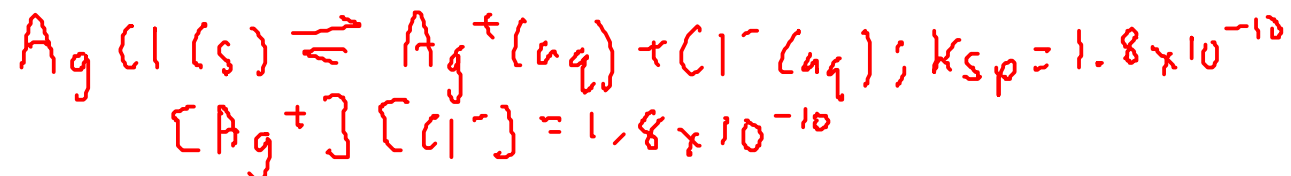


Silver chloride is much less soluble in a solution of 0.1 M NaCl than it is in distilled water. Why? The presence of CHLORIDE ION forces the solubility equilibrium back to the left, meaning less silver chloride can dissolve!

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 (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]	$\Delta$	[Equilibrium]
$\text{Ag}^+$	0	+x	x
$\text{Cl}^-$	0	+x	x

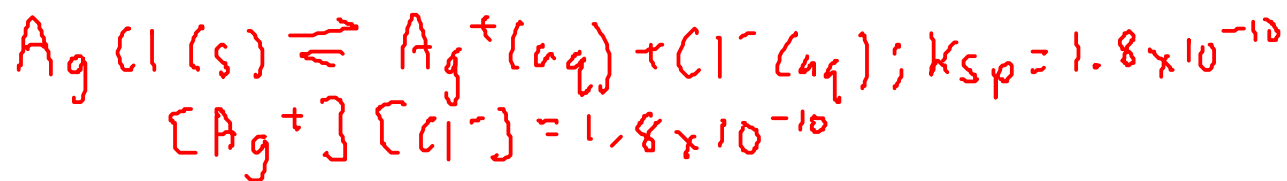
$$(x)(x) = 1.8 \times 10^{-10}$$

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

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

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

1.9 ppm  
AgCl,  
distilled  
H<sub>2</sub>O



For solubility in 0.10 M NaCl solution...

Species	[Initial]	$\Delta$	[Equilibrium]
$\text{Ag}^+$	0	+x	x
$\text{Cl}^-$	0.10	+x	0.10+x

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

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

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

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

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

in 0.10 M NaCl

(0.00026 ppm)

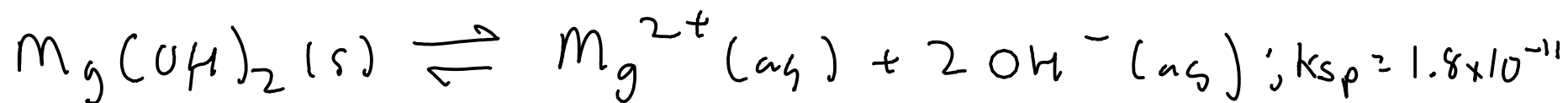
Compare:

1.9 ppm AgCl in distilled water

0.00026 ppm 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