Solubility Equilibria

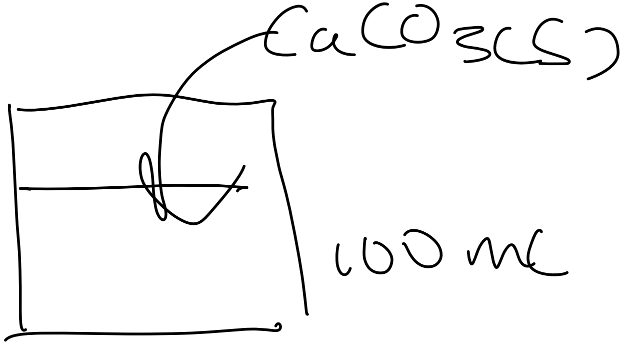
- All ionic compounds dissolve in water to some degree.
  - However, many compounds have such low solubility in water that we classify them as insoluble.

insoluble

$$CaCO_{3(s)} \rightleftharpoons Ca^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$$

Keq defined Ksp

Ksp 10^{-9} reactent favored



$$\frac{1g}{100mL}$$

$$1g\ CaCO_3 \times \frac{1mol}{110g} =$$

$$\frac{9.1\times10^{-3}\ mol}{0.100} = 9.1\times10^{-4}M$$

molar solubility
"how much of the insoluble salt" will dissolve in water

Calculating Ksp from molar solubility

Determine the K_{sp} of silver bromide, given that its molar solubility is 5.71×10^{-7} moles per liter. → at eq.

$$AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq)$$

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

$$AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq)$$

I: 0 and 0

C: $+x$ and $+x$

E: 5.71×10^{-7} and 5.71×10^{-7}

$$K_{sp} = (5.71 \times 10^{-7})^2$$

$$= 3.26 \times 10^{-13}$$

Calculating molar solubility from Ksp

"x"

Calculate the molar solubility of calcium hydroxide if the Ksp is 6.5×10^{-6}

$K_{sp} = 6.5 \times 10^{-6}$

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

I: 0, 0

C: $+x$, $+2x$

E

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

$K_{sp} = (x)(2x)^2$

$\sqrt[3]{K_{sp} = 4x^3}$

$\left(\frac{6.5 \times 10^{-6}}{4}\right)^{1/3} = \left(\frac{4x^3}{4}\right)^{1/3}$

$\boxed{0.012\,M}$

$x = \boxed{0.012\,M} = Ca^{2+}$

$2x = \boxed{0.024\,M = OH^-}$

Common ion
0.100 M *F-(aq)*

$$CaF_2(s) \rightleftharpoons Ca^{2+}(aq) + 2\,F^-(aq)$$

coming from soluble salt

Equilibrium shifts left

What is the molar solubility of CaF_2 in a solution containing 0.100 M NaF?

0.100M F^- is common ion

check [common ion] / [Ksp] ignore change 100

$K_{sp} = 3.9 \times 10^{-11}$

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

	Ca^{2+}	$2F^-$
I	Ø	0.100M
C	+x	+2x
E	x	0.100 + ~~2x~~

$K_{sp} = (x)(0.100)^2$

$3.9 \times 10^{-11} = (0.0100)(x)$

3.9×10^{-9} M

molar solubility is "x" value @ eq.

Precipitation

- $Q = K_{sp}$, the solution is saturated, no precipitation
- $Q < K_{sp}$, the solution is unsaturated, no precipitation
- $Q > K_{sp}$, the solution would be above saturation, the salt above saturation will precipitate.

ChlO_3

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

K_{sp}

1.7×10^{-10}

ChlO_4

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

Calculating Q

and comparing to Ksp

Predicting Precipitation Reactions by Comparing *Q* and K_{sp}

A solution containing lead(II) nitrate is mixed with one containing sodium bromide to form a solution that is 0.0150 M in $Pb(NO_3)_2$ and 0.00350 M in NaBr. Does a precipitate form in the newly mixed solution?

$PbBr_2$

$K_{sp} = 4.0 \times 10^{-5}$

$Pb(NO_3)_{2(aq)} + NaBr_{(aq)} \rightarrow PbBr_{2(s)} + NaNO_{3(aq)}$

$PbBr_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2Br^-_{(aq)}$

$Q < K_{sp}$

$Q = [Pb^{2+}][Br^-]^2$

no ppt forms

Qualitative Analysis

$Q = (0.0150)(0.0035)^2$

$Q = 1.8 \times 10^{-7}$

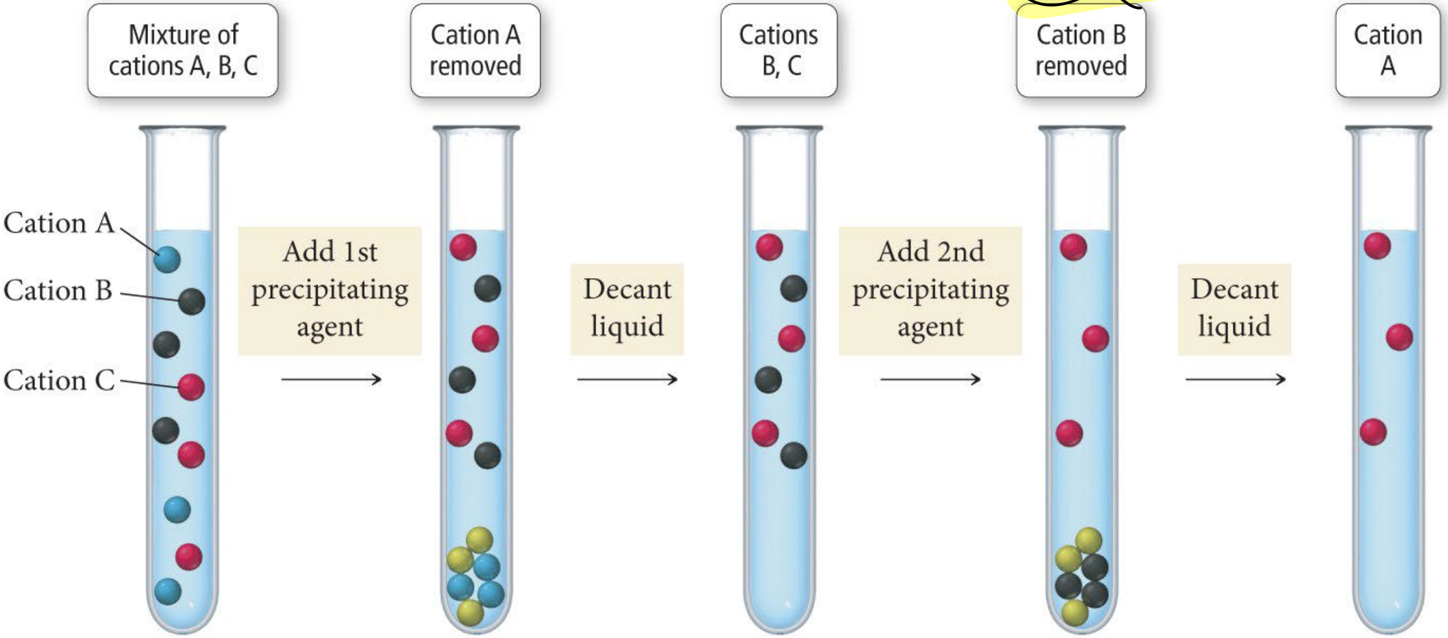
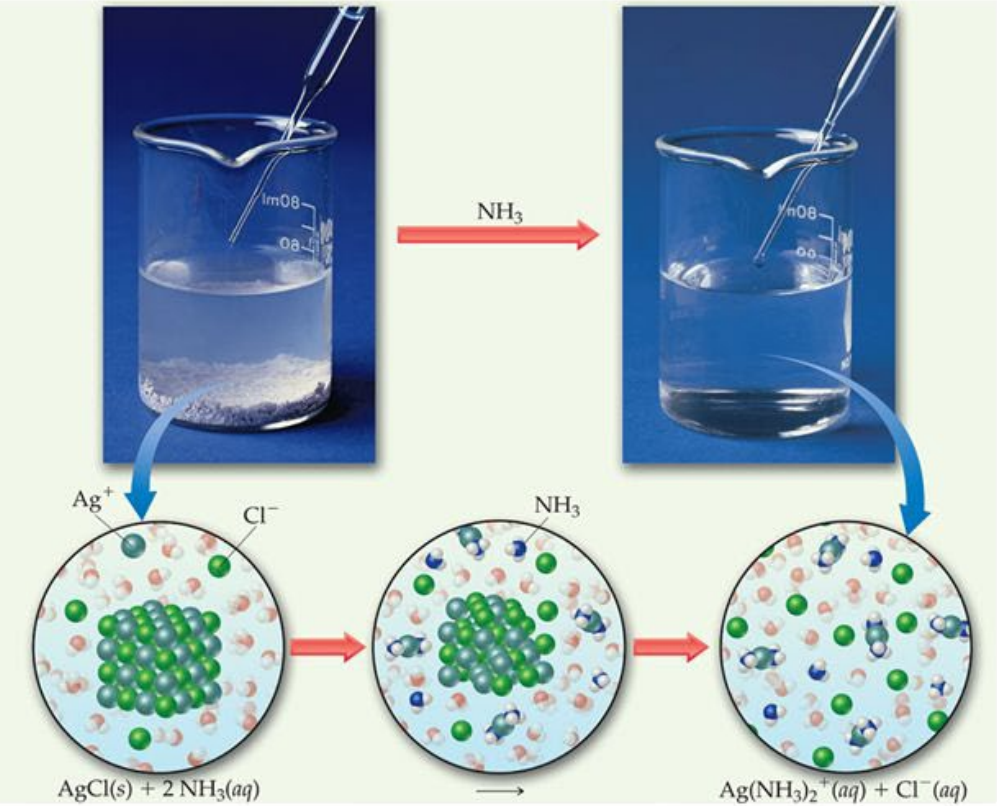


Table 19.4 Formation Constants (K_f) of Some Complex Ions at 25 C

Complex Ion	K_f
$Ag(CN)_2^-$	3.0×10^{20}
$Ag(NH_3)_2^+$	1.7×10^{7}
$Ag(S_2O_3)_2^{3-}$	4.7×10^{13}
AlF_6^{3-}	4×10^{19}
$Al(OH)_4^-$	3×10^{33}
$Be(OH)_4^{2-}$	4×10^{18}
CdI_4^{2-}	1×10^{6}
$Co(OH)_4^{2-}$	5×10^{9}
$Cr(OH)_4^-$	8.0×10^{29}
$Cu(NH_3)_4^{2+}$	5.6×10^{11}
$Fe(CN)_6^{4-}$	3×10^{35}
$Fe(CN)_6^{3-}$	4.0×10^{43}
$Hg(CN)_4^{2-}$	9.3×10^{38}
$Ni(NH_3)_6^{2+}$	2.0×10^{8}
$Pb(OH)_3^-$	8×10^{13}
$Sn(OH)_3^-$	3×10^{25}
$Zn(CN)_4^{2-}$	4.2×10^{19}
$Zn(NH_3)_4^{2+}$	7.8×10^{8}
$Zn(OH)_4^{2-}$	3×10^{15}

Complex Ions



– The formation of these complex ions increases the solubility of these salts.

Other ways to increase solubility of “insoluble” salts