3. 6. Ksp and Solubility

In the last sections, we were able to <u>qualitatively</u> identify which salts were soluble or not soluble.

What about salts that are slightly soluble?

Now, we want to <u>quantify</u> (attach a numerical value) the degree of solubility of the slightly soluble salts.

3. 6. Ksp and Solubility

a) The Solubility Product

i) Salts that are only slightly soluble will form an equilibrium when they dissolve: (don't need much for saturation)

eg. $MgF_{2(s)} \rightleftharpoons Mg^{+2}_{(aq)} + 2F^{-}_{(aq)}$

ii) We can write an equilibrium expression for the solubility of a salt. (solubility product expression)

eg. $Ksp = [Mg^{+2}] [F^{-}]^{2}$

Note: ① attach "sp" to K symbol when dealing with solubility equilibrium.

② It works just like equilibrium expressions! Notice how [MgF$_2$] is not included because it is a solid!

3. 6. Ksp and Solubility

iii) Example: Write the Ksp expression for $Na_2SO_{4(s)}$

$$Na_2SO_{4(s)} \rightleftharpoons 2Na^+_{(aq)} + SO_4^{-2}{}_{(aq)}$$

$$Ksp = [Na^+]^2 [SO_4^{-2}]$$

3. 6. Ksp and Solubility

b) Meaning of Ksp

i) High Ksp value = higher concentration of ions in solution = High Solubility

ii) Low Ksp value = lower concentration of ions in solution = Low Solubility

iii) Ksp is a constant (solubility product constant)

iv) see pg. 333 in Hebden

3. 6. Ksp and Solubility

c) Experimentally Finding Ksp

Method 1: ① Simply take $MgF_{2(s)}$ and add to water until solution is saturated.

② If we know mass of MgF_2 added and water volume we can find $[MgF_2]$ and then we know that:

$[Mg^{+2}] = [MgF_2]$ and $[F^-] = 2 \times [MgF_2]$

③ $Ksp = [Mg^{+2}] [F^-]^2$

Method 2: ① Mix together a source of Mg^{+2} such as $MgSO_{4(aq)}$ and a source of F^- such as $NaF_{(aq)}$.

② Let ppt. of $MgF_{2(s)}$ form and "analyze" solution to find $[Mg^{+2}]$ and $[F^-]$.

3. 6. Ksp and Solubility

d) Ksp Calculations

Type 1 (Find Ksp from ion concentrations)

What is the Ksp for $PbCl_2$ if $[Pb^{+2}]$ is 1.1×10^{-4} M and $[Cl^{-}]$ is 0.33 M?

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$Ksp = [Pb^{+2}][Cl^-]^2$ $Ksp = (1.1 \times 10^{-4}\ M)(0.33\ M)^2$

$Ksp = 1.2 \times 10^{-5}$

3. 6. Ksp and Solubility

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Ksp = $[Ag^{+}]\,[Br^{-}] = (8.8 \times 10^{-7}\ M)^2 =$ **7.7×10^{-13}**

3. 6. Ksp and Solubility

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1.65×10^{-5} M $\quad$ 1.65×10^{-5} M $\quad$ 3.30×10^{-5} M

$Ksp = [Zn^{+2}][OH^{-}]^2 = (1.65 \times 10^{-5} M)(3.30 \times 10^{-5} M)^2 = \mathbf{1.8 \times 10^{-14}}$

3. 6. Ksp and Solubility

Type 2 (Find ion concentrations from Ksp value)

What is the concentration of Ca^{+2} and CO_3^{-2} ions if the Ksp for $CaCO_3$ is 4.8×10^{-9}?

① $CaCO_{3(s)} \rightarrow Ca^{+2}_{(aq)} + CO_3^{-2}{}_{(aq)}$

② Let $[CaCO_3] = X$

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

X X X

④ $Ksp = [Ca^{+2}][CO_3^{-2}]$

⑤ $4.8 \times 10^{-9} = [X][X]$ $4.8 \times 10^{-9} = X^2$

$X = 6.9 \times 10^{-5}$ M **Therefore; $[Ca^{+2}] = [CO_3^{-2}] = 6.9 \times 10^{-5}$ M**

3. 6. Ksp and Solubility

The Ksp for MgF_2 is 6.4×10^{-9}.

a) What is the $[Mg^{+2}]$ and $[F^-]$?

$MgF_{2(s)} \rightleftharpoons Mg^{+2}_{(aq)} + 2F^-_{(aq)}$ $\quad Ksp = [Mg^{+2}] [F^-]^2$

X $\quad$ X $\quad$ 2X

$6.4 \times 10^{-9} = [Mg^{+2}] [F^-]^2$

$6.4 \times 10^{-9} = [X] [2X]^2$

$6.4 \times 10^{-9} = 4X^3$

1.2×10^{-3} M = X Therefore; **$[Mg^{+2}] = 1.2 \times 10^{-3}$ M and $[F^-] = 2.4 \times 10^{-3}$ M**

3. 6. Ksp and Solubility

b) What is the molar solubility of MgF_2?

Molar solubility is mol per litre or M.

Therefore, molar solubility of MgF_2 is simply **1.2×10^{-3} M**

c) What is the solubility of MgF_2 in g/L?

1.2×10^{-3} mol/L x 62.3 g/mol = **7.3×10^{-2} g/L**

Do Questions: #40 page 91; #41 page 92; #42-55 page 95