SOLUBILITY PRODUCT

Solubility

- ionic compounds tend to be **insoluble in non-polar solvents**
- ionic compounds tend to be **soluble in water**
- water is a polar solvent and stabilises the separated ions
- some ionic compounds are very insoluble
- even soluble ionic compounds have a limit as to how much solute dissolves

Saturated solutions

- a solution becomes saturated when solute no longer dissolves in the solvent
- solubility varies with temperature
- most solutes are more soluble at higher temperatures

Solubility product

Even the most insoluble ionic compounds dissolve to a small extent.
An equilibrium is set up between the undissolved solid and its aqueous ions;

eg (i) $\mathbf{MX}(s) \rightleftharpoons \mathbf{M^+}(aq) + \mathbf{X^-}(aq)$

(ii) $\mathbf{BaSO_4}(s) \rightleftharpoons \mathbf{Ba^{2+}}(aq) + \mathbf{SO_4^{2-}}(aq)$

(iii) $\mathbf{PbCl_2}(s) \rightleftharpoons \mathbf{Pb^{2+}}(aq) + \mathbf{2Cl^-}(aq)$

Theory

Applying the equilibrium law to (i) and assuming the concentration of MX(s) is constant in a saturated solution.
[ ] is the concentration in mol dm $^{-3}$

$[M^+(aq)]\,[X^-(aq)]$ = a constant

The constant is known as the SOLUBILITY PRODUCT K_{sp}

$\mathbf{AgCl}(s)$	$\rightleftharpoons$	$\mathbf{Ag^+}(aq)$	+	$\mathbf{Cl^-}(aq)$	K_{sp} =	$[Ag^+(aq)]\,[Cl^-(aq)]$
$\mathbf{BaSO_4}(s)$	$\rightleftharpoons$	$\mathbf{Ba^{2+}}(aq)$	+	$\mathbf{SO_4^{2-}}(aq)$	K_{sp} =	$[Ba^{2+}(aq)]\,[SO_4^{2-}(aq)]$
$\mathbf{PbCl_2}(s)$	$\rightleftharpoons$	$\mathbf{Pb^{2+}}(aq)$	+	$\mathbf{2Cl^-}(aq)$	K_{sp} =	$[Pb^{2+}(aq)]\,[Cl^-(aq)]^2$

Units

The value of K_{sp} has units and is varies with temperature

$K_{sp} = [Ag^+(aq)]\,[Cl^-(aq)]$ units of **$mol^2\,dm^{-6}$**

$K_{sp} = [Pb^{2+}(aq)]\,[Cl^-(aq)]^2$ units of **$mol^3\,dm^{-9}$**

Q.1 *Write expressions for K_{sp} for the following compounds; state the units of K_{sp}*

AgBr

$CaSO_4$

$Fe(OH)_2$

$Fe(OH)_3$

Calculating solubility Solubility products can be used to calculate the solubility of ionic compounds.

Example 1 *At 25°C the solubility product of lead(II) sulphide, PbS is 4 x 10^{-28} mol^2 dm^{-6} . Calculate the solubility of lead(II) sulphide.*

The equation for its solubility is **PbS**(s) $\rightleftharpoons$ **Pb^{2+}**(aq) **+** **S^{2-}**(aq)

The expression for the solubility product is **K_{sp} = [Pb^{2+}(aq)] [S^{2-}(aq)]**

According to the equation you get one **Pb^{2+}**(aq) for every one **S^{2-}**(aq); the concentrations will be equal **[Pb^{2+}(aq)] = [S^{2-}(aq)]**

Substituting and rewriting the expression for K_{sp} **K_{sp} = [Pb^{2+}(aq)]^2**

Re-arranging **[Pb^{2+}(aq)] = $\sqrt{K_{sp}}$ = $\sqrt{4 \times 10^{-28}}$** = **2 x 10^{-14} mol dm^{-3}**

As you get one Pb^{2+} from one PbS, the solubility of PbS = **2 x 10^{-14} mol dm^{-3}**

M_r for PbS is 239; the solubility is **239 x 2 x 10^{-14} g dm^{-3}** = **4.78 x 10^{-12} g dm^{-3}**
[mass = moles x molar mass]

Example 2 *The solubility of ionic compound MY at 25°C is 5 x 10^{-10} g dm^{-3} . The relative mass of MY is 200. Calculate the solubility product of the salt MY at 25°C.*

Solubility of MY $\dfrac{\text{solubility in g}}{\text{molar mass}} = \dfrac{5 \times 10^{-10}}{200} = 2.5 \times 10^{-12}$ mol dm^{-3}

The equation for its solubility is **MY(s) $\rightleftharpoons$ M^+(aq) + Y^-(aq)**

The expression for the solubility product is **K_{sp} = [M^+(aq)] [Y^-(aq)]**

According to the equation; moles of M^+(aq) = moles of Y^-(aq)
= moles of MY(s) which have dissolved

Substituting values K_{sp} = [2.5 x 10^{-12}] [2.5 x 10^{-12}]

The value of the solubility product K_{sp} **= 6.25 x 10^{-24} mol^2 dm^{-6}**

Q.2 a) *Calculate the solubility of $PbSO_4$ in mol dm^{-3} (K_{sp} = 1.6 x 10^{-8} mol^2 dm^{-6}).*

b) *The solubility of Ag_2CO_3 is 1.2 x 10^{-4}mol dm^{-3} ; calculate the K_{sp} value.*

THE COMMON ION EFFECT

Introduction Adding a common ion, (**one which is present in the solution**), will result in the precipitation of a sparingly soluble ionic compound.

eg Adding a solution of sodium chloride to a saturated solution of silver chloride will result in the precipitation of silver chloride.

Theory Silver chloride dissociates in water as follows **AgC*l***(s) $\rightleftharpoons$ **Ag**$^+$(aq) **+ C*l***$^-$(aq)

The expression for the solubility product is **K_{sp} = [Ag$^+$(aq)] [C*l*$^-$(aq)]**

which, at 25°C, is **= 1.2 x 10^{-10} mol^2 dm^{-6}**

If this value is exceeded, precipitation will occur.

The value can be exceeded by adding **EITHER** of the two soluble ions.

If sodium chloride solution is added, the concentration of C*l*$^-$(aq) will increase and precipitation will occur.

Likewise, addition of silver nitrate solution $AgNO_3$(aq) will produce the same effect as it would increase the concentration of Ag^+(aq).

Example *If equal volumes of the following solutions are mixed, will AgCl be precipitated?*
$AgNO_3$ (2 x 10^{-5} mol dm^{-3}) and NaCl (2 x 10^{-5} mol dm^{-3}). K_{sp} = 1.2 x10^{-10} mol^2 dm^{-6}

- *in $AgNO_3$ the concentration of Ag^+ is 2 x 10^{-5} mol dm^{-3}*
  in NaCl the concentration of Cl^- is 2 x 10^{-5} mol dm^{-3}
- *when equal volumes are mixed, the concentrations are halved*
  [Ag^+] = 1 x 10^{-5} mol dm^{-3} [Cl^-] = 1 x 10^{-5} mol dm^{-3}
- *[Ag^+] [Cl^-] = [1 x 10^{-5} mol dm^{-3}] x [1 x 10^{-5} mol dm^{-3}] = 1 x 10^{-10} mol^2 dm^{-6}*
- *because this is lower than the K_{sp} for AgCl...* ***NO PRECIPITATION OCCURS***

Q.3 *Which ion(s) would you add to form a precipitate in the following saturated solutions?*

a) AgI *b) $BaSO_4$* *c) PbS*

Q.4 *Describe what would happen (and explain why it happens) when...*

a) Dilute hydrochloric acid is added to a saturated solution of silver(I) chloride.

b) Sodium sulphate is added to a saturated solution of barium sulphate.