

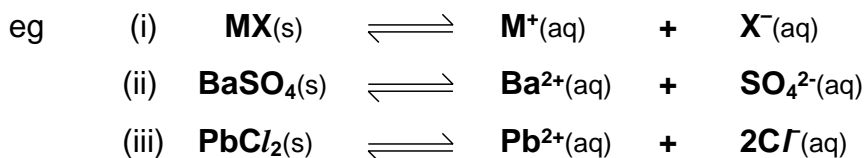
## SOLUBILITY PRODUCT

- 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**

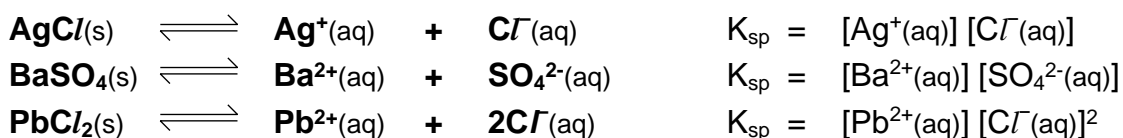
Even the most insoluble ionic compounds dissolve to a small extent. An equilibrium is set up between the undissolved solid and its aqueous ions;



**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<sup>-3</sup>

$$[\mathbf{M^+(aq)}] [\mathbf{X^-(aq)}] = \text{a constant}$$

The constant is known as the **SOLUBILITY PRODUCT K<sub>sp</sub>**



**Units** The value of K<sub>sp</sub> **has units** and is **varies with temperature**

$K_{sp} = [\text{Ag}^+(\text{aq})] [\text{Cl}^-(\text{aq})]$	units of	<b>mol<sup>2</sup> dm<sup>-6</sup></b>
$K_{sp} = [\text{Pb}^{2+}(\text{aq})] [\text{Cl}^-(\text{aq})]^2$	units of	<b>mol<sup>3</sup> dm<sup>-9</sup></b>

**Q.1** Write expressions for K<sub>sp</sub> for the following compounds; state the units of K<sub>sp</sub>

AgBr

CaSO<sub>4</sub>

Fe(OH)<sub>2</sub>

Fe(OH)<sub>3</sub>

### 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  $4 \times 10^{-28} \text{ mol}^2 \text{ dm}^{-6}$ . Calculate the solubility of lead(II) sulphide.

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

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

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

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

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

As you get one  $\text{Pb}^{2+}$  from one PbS, the solubility of PbS =  $2 \times 10^{-14} \text{ mol dm}^{-3}$

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

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

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

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

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

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

Substituting values  $K_{\text{sp}} = [2.5 \times 10^{-12}][2.5 \times 10^{-12}]$

The value of the solubility product  $K_{\text{sp}} = 6.25 \times 10^{-24} \text{ mol}^2 \text{ dm}^{-6}$

- Q.2**
- Calculate the solubility of  $\text{PbSO}_4$  in  $\text{mol dm}^{-3}$  ( $K_{\text{sp}} = 1.6 \times 10^{-8} \text{ mol}^2 \text{ dm}^{-6}$ ).
  - The solubility of  $\text{Ag}_2\text{CO}_3$  is  $1.2 \times 10^{-4} \text{ mol dm}^{-3}$ ; calculate the  $K_{\text{sp}}$  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  $\text{AgCl(s)} \rightleftharpoons \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$

The expression for the solubility product is  $K_{sp} = [\text{Ag}^+(\text{aq})] [\text{Cl}^-(\text{aq})]$   
 which, at 25°C, is  $= 1.2 \times 10^{-10} \text{ mol}^2 \text{ dm}^{-6}$

**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  $\text{Cl}^-(\text{aq})$  will increase and precipitation will occur.

Likewise, addition of silver nitrate solution  $\text{AgNO}_3(\text{aq})$  will produce the same effect as it would increase the concentration of  $\text{Ag}^+(\text{aq})$ .

**Example** *If equal volumes of the following solutions are mixed, will AgCl be precipitated?  $\text{AgNO}_3$  ( $2 \times 10^{-5} \text{ mol dm}^{-3}$ ) and  $\text{NaCl}$  ( $2 \times 10^{-5} \text{ mol dm}^{-3}$ ).  $K_{sp} = 1.2 \times 10^{-10} \text{ mol}^2 \text{ dm}^{-6}$*

- in  $\text{AgNO}_3$  the concentration of  $\text{Ag}^+$  is  $2 \times 10^{-5} \text{ mol dm}^{-3}$   
 in  $\text{NaCl}$  the concentration of  $\text{Cl}^-$  is  $2 \times 10^{-5} \text{ mol dm}^{-3}$
- when equal volumes are mixed, the concentrations are halved  
 $[\text{Ag}^+] = 1 \times 10^{-5} \text{ mol dm}^{-3}$        $[\text{Cl}^-] = 1 \times 10^{-5} \text{ mol dm}^{-3}$
- $[\text{Ag}^+][\text{Cl}^-] = [1 \times 10^{-5} \text{ mol dm}^{-3}] \times [1 \times 10^{-5} \text{ mol dm}^{-3}] = 1 \times 10^{-10} \text{ mol}^2 \text{ dm}^{-6}$
- 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)  $\text{AgI}$                                       b)  $\text{BaSO}_4$                                       c)  $\text{PbS}$

**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.