Chem Factsheet

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Solubility and Solubility Product

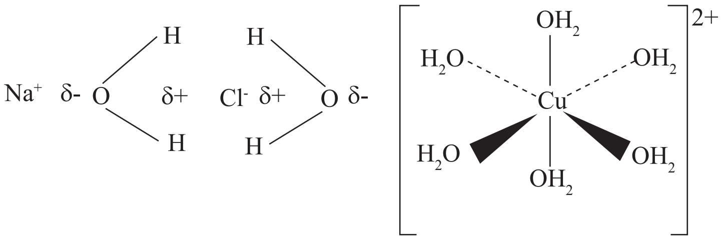
Describing ionic substances as "*soluble in water*" or "*insoluble in water*" is a common occurrence in chemistry.

An ionic substance (the *solute*) is usually described as soluble in water (e.g. sodium chloride) if it mixes homogeneously with water (the *solvent*) at a molecular-ionic level to form a transparent mixture, usually referred to as a *solution*. This is often represented by

$$NaCl(s) + aq \rightarrow Na^+(aq) + Cl^-(aq)$$

This implies that the ions from the solid become individual, separate ions in solution with water molecules actually bonded to each ion. The ions are said to be *hydrated* with the water molecules bonded by ion-dipole attractions or coordinate (dative) bonds as seen in fig. 1.

Fig. 1



Ion-dipole attractions Coordinate bonds

In contrast, an ionic substance is usually described as *insoluble* if it forms a heterogeneous mixture when added to water. This **appears** to consist of the solid ionic substance and colourless water. However, the liquid phase is in fact a **very dilute solution** of the substance containing separate hydrated ions as before.

This is best represented as an equilibrium process because eventually the rate at which ions leave the surface of the solid to enter the solution equals the rate at which ions are deposited from the solution on to the surface of the solid.

$$\text{e.g. } BaSO_4(s) + aq \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$$

This implies that the very dilute solution is in fact also *saturated* - no more can dissolve at the operating temperature.

It is useful to know which ionic substances are soluble or insoluble. There are a few general rules but, in most cases, it is a matter of learning particular cases. The following summary should help for common ionic compounds.

Rule 1	All group 1A compounds, all ammonium compounds and all nitrates are **soluble**.
Rule 2	All chlorides (Cl^-), bromides (Br^-), and iodides (I^-) are **soluble** except for those of Ag^+ and Pb^{2+}.
Rule 3	All sulphates (SO_4^{2-}) are **soluble** except those of Ba^{2+}, Sr^{2+}, Ca^{2+}, and Pb^{2+}.
Rule 4	All hydroxide (OH^-) compounds are i**nsoluble** except those of Group IA, ammonium, Ba^{2+}, Ca^{2+}, and Sr^{2+}.
Rule 5	All carbonates (CO_3^{2-}) are **insoluble** except for those of NH_4^+ and Group IA.
Rule 6	For group 2 sulphates, **solubility decreases** down the group
Rule 7	For group 2 hydroxides, **solubility increases** down the group
Rule 8	For insoluble halides, **solubility decreases** down group 7.

Q1 Use these rules to complete the following table by deciding whether a precipitate should form, or not, when the indicated compounds are mixed in solution. Answer YES or NO.

A	Calcium nitrate and sodium chloride	YES / NO
B	Calcium nitrate and sodium carbonate	YES / NO
C	Barium nitrate and sodium sulphate	YES / NO
D	Barium nitrate and sodium hydroxide	YES / NO

Q2 The nitrate of metal M (M^{p+}) is mixed with the sodium salt of anion A (A^{q-}) in aqueous solution. What can be concluded if (a) a precipitate is formed (b) a precipitate is not formed? Explain.

Solubility

The mixing process (*dissolving*) of a "soluble" substance does not carry on indefinitely. All "soluble" ionic substances eventually reach a limit of solubility when the solution is described as *saturated*.

At this point, the process can also be represented as an equilibrium process (as for the $BaSO_4$ example above) because the solid will be in equilibrium with its saturated solution.

$$NaCl(s) + aq \rightleftharpoons Na^+(aq,sat) + Cl^-(aq,sat)$$

The only difference between water *soluble* and *insoluble* substances is therefore the amount that must dissolve to form a saturated solution. The amount will be relatively high for a soluble substance but relatively low for an insoluble substance. Hence, the use of other phrases such as *sparingly soluble* to describe those substances that are usually considered to be insoluble.

The amount of substance dissolved in a saturated solution (its *solubility*) depends on the actual substance, the pressure and the temperature. In general, solubilities of ionic substances increase with increasing temperature but are usually quoted at 25°C (298K). Pressures have negligible effects but are still quoted at 100kPa.

Definition : The solubility of a substance is the *maximum* amount (i.e. saturated) (usually measured in g or moles) of that substance that may be dissolved in a specified amount of water (usually 100 or 1000 cm^3) at 100kPa and a specified temperature (usually 298K).

e.g. (a) Solubility of NaCl = 359 g per 1000cm^3 at 298K

(b) Solubility of $BaSO_4$ = 0.0002448 g per 100cm^3 at 298K

Q3 Convert each of these solubilities to mol dm^{-3}.

Solubility Products

These are K_c type equilibrium constants applied to the equilibrium described earlier for a sparingly soluble solid in equilibrium with its component aqueous ions at constant temperature.

Note: The following ideas **do not** apply to "soluble" substances (e.g. NaCl) because the ion density in such solutions is so high that interactions between ions make the equilibrium law unreliable.

For a general sparingly soluble ionic substance, M_pA_q, the equilibrium at constant temperature is: $M_pA_q(s) \rightleftharpoons pM^{q+}(aq) + qA^{p-}(aq)$

Applying the equilibrium law gives: $Kc = \dfrac{[M^{q+}(aq)]^p \times [A^{p-}(aq)]^q}{[M_pA_q(s)]}$

where [X] represents the concentration of substance X measured in mol dm^{-3}.

However, the concentration of the solid phase is constant (C). The K_c expression can therefore be rewritten as: $K_c \times C = [M^{q+}(aq)]^p \times [A^{p-}(aq)]^q$. Kc × C is the product of 2 constants which is therefore another constant. This combined constant is labelled $\mathbf{K_{sp}}$ and called the **solubility product** of M_pA_q.

$$\text{Hence, } K_{sp} = [M^{q+}(aq)]^p \times [A^{p-}(aq)]^q$$

Note: Like ionic product of water (K_w), the word "product" in solubility product refers to the mathematical term, not the chemical term. Mathematically, product means the result of multiplication. Hence, the multiplication of the ion concentrations in the K_{sp} expression.

Note: The **units** of K_{sp} are given by (mol dm^{-3})$^{p+q}$

Q4 For each of the following sparingly soluble substances

1. AgBr 2. $PbCl_2$ 3. $Fe(OH)_3$ 4. $Ca_3(PO_4)_2$

(a) write the equation for the equilibrium present in a saturated solution

(b) write the expression for K_{sp}

(c) deduce the units of K_{sp}

By their very nature, sparingly soluble substances produce saturated solutions with low to extremely low ion concentrations. Hence, when these are substituted into a K_{sp} expression, the values of K_{sp} are generally very low. This and variations can be seen in the data shown in table 1.

Table 1

Fluorides	K_{sp}	Iodides	K_{sp}	Hydroxides	K_{sp}
MgF_2	7.4×10^{-11}	PbI_2	8.5×10^{-9}	$Mg(OH)_2$	5.6×10^{-12}
CaF_2	1.5×10^{-10}	AgI	8.5×10^{-17}	$Ca(OH)_2$	4.7×10^{-6}
SrF_2	2.5×10^{-9}	Carbonates	K_{sp}	$Sr(OH)_2$	6.4×10^{-3}
BaF_2	1.8×10^{-7}	$MgCO_3$	6.8×10^{-6}	$Ba(OH)_2$	5.0×10^{-3}
Chlorides	K_{sp}	$CaCO_3$	5.0×10^{-9}	$Cu(OH)_2$	1.6×10^{-19}
$PbCl_2$	1.2×10^{-5}	$SrCO_3$	5.6×10^{-10}	$Cr(OH)_3$	6.7×10^{-31}
AgCl	1.9×10^{-10}	$CuCO_3$	2.5×10^{-10}	Sulphates	K_{sp}
Bromides	K_{sp}	$FeCO_3$	2.1×10^{-11}	$CaSO_4$	7.1×10^{-5}
$PbBr_2$	6.6×10^{-6}	Phosphates	K_{sp}	$SrSO_4$	3.5×10^{-7}
AgBr	5.4×10^{-13}	$Ca_3(PO_4)_2$	2.0×10^{-29}	$BaSO_4$	1.1×10^{-10}

Q5 Use the data in table 1 to complete each of the following statements about solubility trends

(a) Down group 2, sulphate solubilities DECREASE / INCREASE
(b) Down group 2, hydroxide solubilities DECREASE / INCREASE
(c) Down group 2, fluoride solubilities DECREASE / INCREASE
(d) Down group 2, carbonate solubilities DECREASE / INCREASE
(e) Down group 7, halide solubilities DECREASE / INCREASE

The significance of K_{sp} values

The values of K_{sp} define the conditions necessary for a precipitate of a sparingly soluble substance to form when solutions containing the metal ion and the anion are mixed.

If $[M^{q+}(aq)]$ and $[A^{p-}(aq)]$ in **any** solution are substituted into the K_{sp} expression ($[M^{q+}(aq)]^p \times [A^{p-}(aq)]^q$) and the value derived is:

1. Greater than K_{sp}, a precipitate will form
2. Less than or equal to K_{sp}, a precipitate will not form.

Q6 Using the K_{sp} data table, will a precipitate of lead chloride form when 10.0 cm^3 of 0.05 mol dm^{-3} lead nitrate is mixed with 10.0 cm^3 0.02 mol dm^{-3} sodium chloride at 298K?

Q7 Using the K_{sp} data table, will a precipitate of silver chloride form when one drop (assume this is 0.05 cm^3) of 0.10 mol dm^{-3} silver nitrate is mixed with 10.0 cm^3 0.10 mol dm^{-3} sodium chloride at 298K?

Measuring solubility products

This involves measuring the concentration of one of the component ions in a saturated solution of the salt at constant temperature. The concentration of the other ion can then be deduced from the formula of the salt.

e.g.1 For $PbCl_2$, if $[Pb^{2+}]$ is measured at "x" mol dm^{-3}, then $[Cl^-] = 2x$
e.g.2 For Ag_2CrO_4, if $[Ag^+]$ is measured at "x" mol dm^{-3}, then $[CrO_4^{2-}] = x/2$.

Q8 (a) For $Fe(OH)_3$, if $[Fe^{3+}]$ is measured at "x" mol dm^{-3}, what is $[OH^-]$?
(b) For $Al(OH)_3$, if $[OH^-]$ is measured at "x" mol dm^{-3}, what is $[Al^{3+}]$?

The problem with making such measurements is that the concentration to be measured is very low. Hence, very sensitive techniques need to be employed.

Two common methods are colorimetry and electrode potential measurements.

If one of the ions is coloured (e.g. Cu^{2+} or CrO_4^{2-}), comparison of the absorbance by the saturated solution with those of an appropriate range of standard solutions can give the concentration of the coloured ion.

Note: See FactSheet 329 for further details.

Alternatively, electrode potential measurements make use of the Nernst equation to measure the concentration of the metal ion (M^{q+}) in the saturated solution. The general equation is:

$$E_M = E^\ominus_M + \frac{0.059}{q}\log_{10}[M^{q+}(aq)]$$

For example if the potential of a silver electrode dipping into a saturated solution of silver chloride relative to a standard hydrogen electrode at 298K is measured (i.e $E_{Ag} = +0.512V$) then using $E^\ominus_{Ag} = 0.799V$ and q = 1 for the Ag^+ ion, gives $[Ag^+(aq)] = 1.37 \times 10^{-5}$ mol dm^{-3}.

Note: See FactSheet 199 for full practical details.

Since $Ag^+ : Cl^- = 1 : 1$, $[Cl^-(aq)]$ is also 1.37×10^{-5} mol dm^{-3}.
$\rightarrow K_{sp}[AgCl] = 1.37 \times 10^{-5} \times 1.37 \times 10^{-5} = 1.88 \times 10^{-10}$ $mol^2\ dm^{-6}$.

Q9 The potential of a copper electrode dipping into a saturated solution of copper hydroxide relative to a standard hydrogen electrode at 298K is +0.149V. Using $E^\ominus_{Cu} = +0.340V$ calculate the solubility product of copper(II) hydroxide at 298K.

Converting solubility to solubility product

e.g.1 The solubility of magnesium hydroxide is 0.0065 g dm^{-3} at 298K. Calculate K_{sp} of magnesium hydroxide at 298K.
$[Mg(OH)_2(aq)] = 0.0065$ g dm^{-3}
$= 0.0065 / (24.1 + 2(16+1))$ mol dm^{-3}
$= 1.119 \times 10^{-4}$ mol dm^{-3}
$Mg(OH)_2(s) + aq \rightarrow Mg(OH)_2(aq) \rightarrow Mg^{2+}(aq) + 2OH^-(aq)$
$\rightarrow [Mg^{2+}(aq)] = 1.119 \times 10^{-4}$ mol dm^{-3}
and $[OH^-(aq)] = 2 \times 1.119 \times 10^{-4} = 2.238 \times 10^{-4}$ mol dm^{-3}
$\rightarrow K_{sp} = [Mg^{2+}(aq)]\,[OH^-(aq)]^2$
$= 1.119 \times 10^{-4} \times (2.238 \times 10^{-4})^2$
$= 5.57 \times 10^{-12}$ $mol^3\ dm^{-9}$

e.g.2 The solubility of calcium phosphate is 7.14 x 10^{-7} mol dm^{-3} at 298K. Calculate K_{sp} of calcium phosphate at 298K.
$Ca_3(PO_4)_2(s) + aq \rightarrow Ca_3(PO_4)_2(aq) \rightarrow \mathbf{3}Ca^{2+}(aq) + \mathbf{2}PO_4^{3-}(aq)$
$[Ca_3(PO_4)_2(aq)] = 7.14 \times 10^{-7}$
$\rightarrow [Ca^{2+}(aq)] = \mathbf{3} \times 7.14 \text{ x } 10^{-7} = 2.142 \times 10^{-6}$ mol dm^{-3}
and $[PO_4^{3-}(aq)] = \mathbf{2} \times 7.14 \text{ x } 10^{-7} = 1.428 \times 10^{-6}$ mol dm^{-3}
$\rightarrow K_{sp} = [Ca^{2+}(aq)]^3\,[PO_4^{3-}(aq)]^2$
$= (2.142 \times 10^{-6})^3 \times (1.428 \times 10^{-6})^2$
$= 2.00 \times 10^{-29}$ $mol^5\ dm^{-15}$

Q10 The solubility of chromium hydroxide is 1.293×10^{-6} g dm^{-3} at 298K. Calculate K_{sp} of chromium hydroxide at 298K.

Converting solubility product to solubility

e.g.1 Calculate the solubility of silver bromide at 298K given its K_{sp} is 5.40×10^{-13} $mol^2\ dm^{-6}$ at 298K.
Let solubility AgBr = S mol dm^{-3}
$AgBr(s) + aq \rightarrow AgBr(aq) \rightarrow Ag^+(aq) + Br^-(aq)$
$\rightarrow [Ag^+(aq)] = S$ and $[Br^-(aq)] = S$
$\rightarrow K_{sp} = [Ag^+(aq)] \times [Br^-(aq)] = S \times S = 5.40 \times 10^{-13}$
$\rightarrow S = \sqrt{(5.4 \times 10^{-13})} = \mathbf{7.35 \times 10^{-7}\ mol\ dm^{-3}}$
$= 7.35 \times 10^{-7} \times (107.9 + 79.9) = \mathbf{1.38 \times 10^{-4}\ g\ dm^{-3}}$

e.g.2 Calculate the solubility of lead iodide at 298K given its K_{sp} is 8.50×10^{-9} $mol^3\ dm^{-9}$ at 298K.
Let solubility PbI_2 = S mol dm^{-3}
$PbI_2(s) + aq \rightarrow PbI_2(aq) \rightarrow Pb^{2+}(aq) + 2I^-(aq)$
$\rightarrow [Pb^{2+}(aq)] = S$ and $[I^-(aq)] = 2S$
$\rightarrow K_{sp} = [Pb^{2+}(aq)] \times [I^-(aq)]^2 = S \times (2S)^2 = 4S^3 = 8.50 \times 10^{-9}$
$\rightarrow S = \sqrt[3]{(8.50 \times 10^{-9}/4)} = \mathbf{1.29 \times 10^{-3}\ mol\ dm^{-3}}$
$= 1.29 \times 10^{-3} \times (207.2 + 126.9) = \mathbf{0.431\ g\ dm^{-3}}$

Q11 Calculate the solubility of aluminium hydroxide at 298K given its K_{sp} is 3.12×10^{-34} $mol^3\ dm^{-9}$ at 298K.

The common ion effect

This refers to the reduction in the solubility of an ionic precipitate when a soluble compound containing one of the ions of the precipitate is added to the solution in equilibrium with the precipitate.

For example, silver chloride is less soluble in sodium chloride solution than in water because of the presence of the common chloride ion.

It can be applied quantitatively for sparingly soluble substances where K_{sp} is applicable or qualitatively to these and more soluble substances.

e.g. In a previous example, the solubility of silver bromide at 298K was calculated to be 7.35×10^{-7} mol dm^{-3} using its K_{sp} as 5.40×10^{-13} mol^2 dm^{-6} at 298K. Calculate the solubility of silver bromide in 0.100 mol dm^{-3} sodium chloride solution.

Let the new solubility AgBr = S mol dm^{-3}
$AgBr(s) + aq \rightarrow AgBr(aq) \rightarrow Ag^+(aq) + Br^-(aq)$
Dissolution of the AgBr produces equal concentrations (S) of Ag^+ and Br^- ions but the sodium bromide adds 0.100 mol dm^{-3} of Br^-.
$\rightarrow [Ag^+(aq)] = S$ and $[Br^-(aq)] = S + 0.100$
$\rightarrow K_{sp} = [Ag^+(aq)] \times [Br^-(aq)] = S \times (S + 0.100) = 5.40 \times 10^{-13}$
But $S << 0.100$
$\rightarrow S \times 0.100 \approx 5.40 \times 10^{-13}$
S = **5.40×10^{-12} mol dm^{-3}**

This shows that the solubility of AgBr has been reduced to almost one millionth by the presence of the common Br^- ion at a relatively low concentration.

Qualitatively this can be predicted using Le Chatelier's principle.

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

Addition of bromide ions from the sodium bromide causes the equilibrium to shift to the left in order to reduce the effect of the addition. Hence, silver ions are removed by precipitation of more silver bromide which amounts to the reduction in solubility of the silver bromide.

There are many practical applications of this effect. For example used in areas where drinking water is derived from chalk or limestone rocks. Soluble sodium carbonate is added to the water to reduce the calcium carbonate content and hence, the hardness of the water.

$$CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$$

The common carbonate ion pushes the equilibrium to the left reducing the amount of dissolved calcium ions which is responsible for the hardness. The calcium ions are said to be "salted out".

Q12 The salting out process is also used in the manufacture of soaps which are sodium salts of fatty acids (e.g. $CH_3(CH_2)_{16}COO^-Na^+$) produced by the alkaline hydrolysis of a fat or oil. Using the common ion effect, explain why sodium chloride solution is added to the reaction mixture after hydrolysis.

Answers to Questions

Q1	A → $CaCl_2$ and $NaNO_3$ possible – both soluble	NO
	B → $CaCO_3$ and $NaNO_3$ possible – $CaCO_3$ insoluble	YES
	C → $BaSO_4$ and $NaNO_3$ possible – $BaSO_4$ insoluble	YES
	D → $Ba(OH)_2$ and $NaNO_3$ possible – both soluble	NO

Q2 (a) The compound of M^{p+} with A^{q-} (i.e. M_qA_p) is insoluble
(b) The compound of M^{p+} with A^{q-} (i.e. M_qA_p) is soluble.
Each mixture will contain Na^+, NO_3^-, M^{n+} and A^{m-} ions. Possible compounds from this mixture must involve a positive ion combining with a negative ion. According to rule 1, $NaNO_3$ is always soluble. Hence the experiment tests the solubility of the compound formed from M^{n+} and A^{m-}. A precipitate will not form if this compound is soluble but a precipitate will form if this compound is insoluble.

Q3 (a) Solubility of NaCl = 359 g per 1000cm^3
= 359 / (23.0 + 35.5) = 6.14 mol dm^{-3}
(b) Solubility of $BaSO_4$ = 0.0002448 g per 100cm^3
= 0.0002448 / 233.4 moles per 100cm^3
= (0.0002448 / 233.4) × 1000/100 moles per 1000cm^3
= 1.05×10^{-5} mol dm^{-3}

Q4 1. (a) $AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq)$
(b) $K_{sp} = [Ag^+(aq)]\,[Br^-(aq)]$
(c) Units of K_{sp} = (mol dm^{-3})$^{1+1}$ = mol^2 dm^{-6}
2. (a) $PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$
(b) $K_{sp} = [Pb^{2+}(aq)]\,[Cl^-(aq)]^2$
(c) Units of K_{sp} = (mol dm^{-3})$^{1+2}$ = mol^3 dm^{-9}
3. (a) $Fe(OH)_3(s) \rightleftharpoons Fe^{3+}(aq) + 3OH^-(aq)$
(b) $K_{sp} = [Fe^{3+}(aq)]\,[OH^-(aq)]^3$
(c) Units of K_{sp} = (mol dm^{-3})$^{1+3}$ = mol^4 dm^{-12}
4. (a) $Ca_3(PO_4)_2(s) \rightleftharpoons 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)$
(b) $K_{sp} = [Ca^{2+}(aq)]^3\,[PO_4^{3-}(aq)]^2$
(c) Units of K_{sp} = (mol dm^{-3})$^{3+2}$ = mol^5 dm^{-15}

Q5 (a) -- DECREASE (b) -- INCREASE (c) -- INCREASE
(d) -- DECREASE (e) DECREASE

Q6 As a result of dilution during mixing, $[Pb^{2+}]$ after mixing
= 0.05 × 10/20 = 0.025 mol dm^{-3}
As a result of dilution during mixing, $[Cl^-]$ after mixing
= 0.02 × 10/20 = 0.01 mol dm^{-3}
$K_{sp} = [Pb^{2+}(aq)][Cl^-(aq)]^2 = 1.2 \times 10^{-5}$ mol^3 dm^{-9}
In the mixture, $[Pb^{2+}][Cl^-]^2 = 0.025 \times (0.01)^2 = 2.5 \times 10^{-6}$
This is less than K_{sp} which means a precipitate will not form.

Q7 As a result of dilution during mixing, $[Ag^+]$ after mixing
= 0.10 × 0.05/10.05 = 4.98×10^{-4} mol dm^{-3}
As a result of dilution during mixing, $[Cl^-]$ after mixing
= 0.10 × 10/10.05 = 9.95×10^{-2} mol dm^{-3}
$K_{sp} = [Ag^+(aq)][Cl^-(aq)] = 1.9 \times 10^{-10}$ mol^3 dm^{-9}
In the mixture, $[Ag^+][Cl^-] = 4.98 \times 10^{-4} \times 9.95 \times 10^{-2} = 4.95 \times 10^{-5}$
This is greater than K_{sp} which means a precipitate will form. Precipitation of AgCl will occur until the remaining $[Ag^+]$ and $[Cl^-]$ values satisfy the K_{sp} value.

Q8 (a) $[OH^-] = 3x$ (b) $[Al^{3+}] = x/3$

Q9 Substituting into Nernst equation gives
$0.149 = 0.340 + (0.059/2)\log_{10}[Cu^{2+}]$
$\rightarrow [Cu^{2+}] = 3.353 \times 10^{-7}$ mol dm^{-3}
Also $[OH^-] = 2 \times [Cu^{2+}] = 2 \times 3.353 \times 10^{-7}$
= 6.706×10^{-7} mol dm^{-3}
$\rightarrow K_{sp} = [Cu^{2+}][OH^-]^2 = 3.353 \times 10^{-7} \times (6.706 \times 10^{-7})^2$
= 1.51×10^{-19} mol^3 dm^{-9}

Q10 $[Cr(OH)_3] = [Cr^{3+}] = 1.293 \times 10^{-6} / (52 + 3(16+1)) = 1.255 \times 10^{-8}$ mol dm^{-3}
$Cr(OH)_3 \rightarrow Cr^{3+} + 3OH^-$
$\rightarrow [OH^-] = 3[Cr^{3+}] = 3 \times 1.255 \times 10^{-8} = 3.765 \times 10^{-8}$
$\rightarrow K_{sp} = [Cr^{3+}] \times [OH^-]^3 = 1.255 \times 10^{-8} \times (3.765 \times 10^{-8})^3$
= 6.698×10^{-31} mol^4 dm^{-12}

Q11 Let solubility $Al(OH)_3$ = S mol dm^{-3}
$Al(OH)_3(s) + aq \rightarrow Al(OH)_3(aq) \rightarrow Al^{3+}(aq) + 3OH^-(aq)$
$\rightarrow [Al^{3+}(aq)] = S$ and $[OH^-(aq)] = 3S$
$\rightarrow K_{sp} = [Al^{3+}(aq)] \times [OH^-(aq)]^3 = S \times (3S)^3 = 27S^4 = 3.12 \times 10^{-34}$
$\rightarrow S = \sqrt[4]{(3.12 \times 10^{-34}/27)}$ = **1.84×10^{-9} mol dm^{-3}**
= $1.84 \times 10^{-9} \times (27.0 + 3(16+1))$ = **1.44×10^{-7} g dm^{-3}**

Q12 Addition of sodium chloride reduces the solubility of the soaps and allows them to be separated more easily. The soaps precipitate due to the presence of the common Na^+ ion.

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