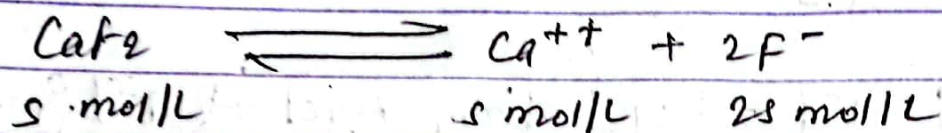


Q.8 The solubility of CaF_2 in water at 18°C is $2.05 \times 10^{-4} \text{ mol l}^{-1}$. Calculate its solubility product.

$$\text{solubility of } \text{CaF}_2 (\text{s}) = 2.05 \times 10^{-4} \text{ mol l}^{-1}$$

$$\text{solubility product of } \text{CaF}_2 (K_{sp}) = ?$$



$$K_{sp} = [\text{Ca}^{++}] [\text{F}^-]^2$$

$$= \text{s} \times (2\text{s})^2$$

$$= 4\text{s}^3$$

$$= 4 \times (2.05 \times 10^{-4})^3$$

$$= 3.346 \times 10^{-11} \text{ mol/l}$$

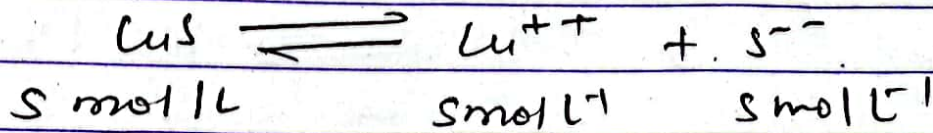
Hence, the solubility of CaF_2 is $3.44 \times 10^{-11} \text{ mol/l}$.

Q. The solubility product of CuS is 8.0×10^{-45} at a certain temperature. Find its solubility.

$$\text{Solubility product of } \text{CuS} (K_{sp}) = 8.0 \times 10^{-45}$$

$$\text{Solubility of } \text{CuS} (\text{s}) = ?$$

CuS ionizes as,



$$K_{sp} = [\text{Cu}^{++}] [\text{S}^{--}]$$

$$K_{sp} = s \times s$$

$$s = \sqrt{K_{sp}}$$

$$s = \sqrt{8.0 \times 10^{-45}}$$

$$s = 8.94 \times 10^{23} \text{ mol/L}^{-1}$$

Hence, the solubility of CuS is $8.94 \times 10^{23} \text{ mol/L}^{-1}$

The solubility of AgCl in water at 298 K is $1.43 \times 10^{-3} \text{ g/L}$. Calculate its solubility in 0.5 M KCl solution.

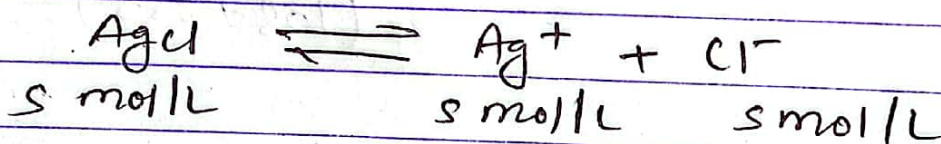
Solution:-

$$\text{Solubility of AgCl in water in mol/L} = \frac{\text{g/L}}{\text{Mol. wt}}$$

$$= \frac{1.43 \times 10^{-3}}{(108 + 35.5)}$$

$$= 9.965 \times 10^{-6} \text{ mol/L.}$$

AgCl ionizes as,

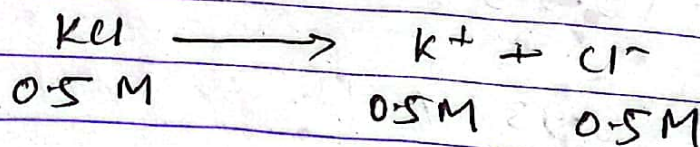


$$\begin{aligned} K_{sp} &= [\text{Ag}^+][\text{Cl}^-] \\ &= s \times s \\ &= (9.965 \times 10^{-6})^2 \end{aligned}$$

$$\therefore K_{sp} \text{ of AgCl in water} = 9.93 \times 10^{-11}$$

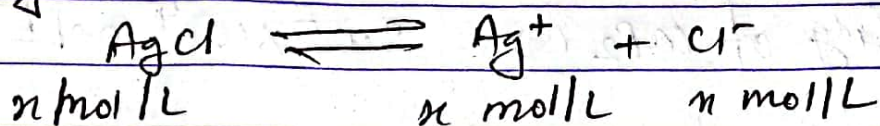
Now,

KCl ionizes as,



let x mol/L be the solubility of AgCl in 0.5 M KCl.

AgCl ionizes as,



$$[\text{Ag}^+] = x \text{ mol/L}$$

& $[\text{Cl}^-] = [\text{Cl}^-] \text{ from AgCl} + [\text{Cl}^-] \text{ from KCl.}$

$$= (x + 0.5) \text{ mol/L.}$$

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

$$= x \times (x + 0.5)$$

x is very small i.e. $x \ll 0.5$. So, $x + 0.5 = 0.5$

$$\therefore K_{sp} = 0.5 \times x$$

$$9.93 \times 10^{-11} = 0.5x \quad [\because K_{sp} \text{ remains same in any soln.}]$$

$$x = \frac{9.93 \times 10^{-11}}{0.5}$$

$$x = 1.986 \times 10^{-10} \text{ mol/L}$$

Hence, the required solubility in 0.5 M KCl is $1.986 \times 10^{-10} \text{ mol/L}$.

Q.

0.00143 gm of AgCl dissolves in one litre of water at 25°C to form a ~~sat~~ saturated solution. What is the solubility product of the salt ($\text{Ag} = 108$, $\text{Cl} = 35.5$)

Given

$$\text{concentration of AgCl sol}^n = 0.00143 \text{ g/L}$$

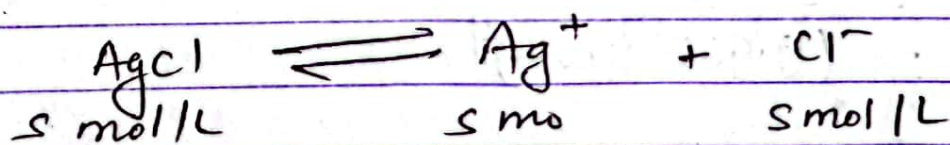
$$K_{sp} \text{ of AgCl} = ?$$

we have,

$$\begin{aligned} \text{Solubility of AgCl(s)} &= \frac{\text{g/L}}{\text{Mol. wt}} \\ &= \frac{0.00143}{(108 + 35.5)} \end{aligned}$$

$$= 9.965 \times 10^{-6} \text{ mol/L}$$

AgCl ionises as,



$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = s^2$$

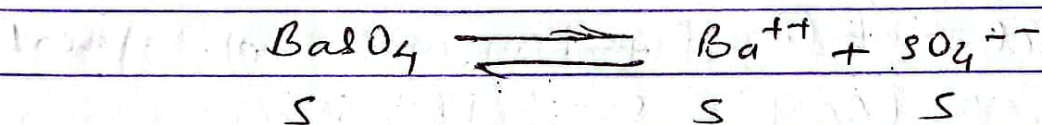
$$= (9.965 \times 10^{-6})^2$$

$$= 9.93 \times 10^{-11}$$

The solubility constant of BaSO_4 in water at 25°C is $1 \times 10^{-10} \text{ mol}^2 \text{L}^{-2}$. Calculate the solubility of BaSO_4 in g/lit ($\text{Ba} = 137$) solution.

$$K_{sp} = 1 \times 10^{-10} \text{ mol}^2 \text{L}^{-2}$$

Solubility of BaSO_4 in mole/lit = ?
 BaSO_4 ionize as,



$$[\text{Ba}^{++}] = s, [\text{SO}_4^{--}] = s$$

$$K_{sp} = [\text{Ba}^{++}] [\text{SO}_4^{--}]$$

$$s \cdot s = 10^{-10}$$

$$s = 10^{-5} \text{ mol/lit}$$

$$\therefore \text{Solubility of } \text{BaSO}_4 = 1 \times 10^{-5} \text{ mol/lit}$$

$$\begin{aligned} \text{Molecular weight of } \text{BaSO}_4 &= 137 + 32 + 4 \times 16 \\ &= 233. \end{aligned}$$

$$\begin{aligned} \text{Solubility in g/lit} &= \text{mol. wt} \times \text{molarity} \\ &= 233 \times 1 \times 10^{-5} \end{aligned}$$

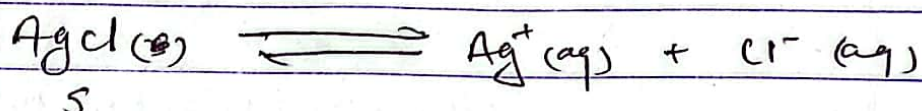
$$= 2.33 \times 10^{-3} \text{ g/l.}$$

Calculate solubility of $\text{AgCl}(s)$ in (a) pure water (b) a solution of 0.1M NaCl at 25°C .

$$K_{sp} \text{ of } \text{AgCl} = 1.0 \times 10^{-10}.$$

Solution

(a) We have,



$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

Let solubility of AgCl in water be ' s ' mol L^{-1}

$$\text{So, } [\text{Ag}^+] = s \quad [\text{Cl}^-] = s.$$

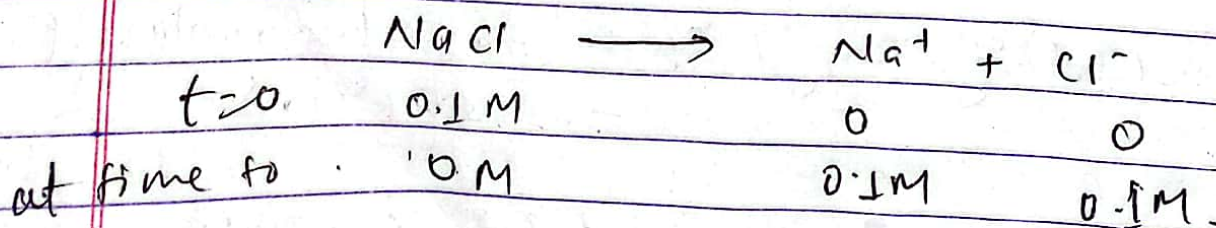
$$K_{sp} = s \times s = s^2$$

$$s = \sqrt{K_{sp}}$$

$$s = \sqrt{1.0 \times 10^{-10}} = 1 \times 10^{-5} \text{ mol L}^{-1}$$

Thus, the solubility of AgCl in pure water
 $= 1 \times 10^{-5} \text{ mol L}^{-1}$

(b) In solution, NaCl is completely ionized



$$[\text{Cl}^-] = 0.1\text{M}$$

Solubility of AgCl in presence of 0.1 M NaCl is 0.1 mol L^{-1}

$$[\text{Ag}^+] = s \text{ mol L}^{-1}$$

$$[\text{Cl}^-] = [\text{Cl}^-]_{\text{AgCl}} + [\text{Cl}^-]_{\text{NaCl}}$$

$$s + 0.1$$

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

$$1 \times 10^{-10} = s(s + 0.1)$$

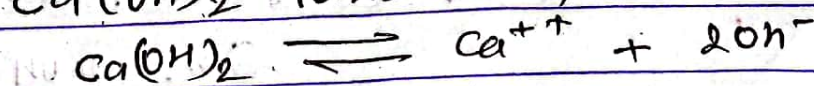
$$= s(0.1) \quad [\text{since } s \ll 0.1 \text{ so } s + 0.1 \approx 0.1]$$

$$s = \frac{10^{-10}}{0.1} = 1 \times 10^{-9} \text{ M}$$

\therefore Thus, the solubility of AgCl in 0.1 M NaCl is, $1 \times 10^{-9} \text{ mol L}^{-1}$.

The solubility product (K_{sp}) of Ca(OH)_2 at 25°C is 4.42×10^{-5} . A 500 ml of a saturated solution of Ca(OH)_2 is mixed with an equal volume of 0.4 M NaOH . How much Ca(OH)_2 is precipitated?

Ca(OH)_2 ionizes as,



$$s \text{ mol/L} \quad s \text{ mol/L} \quad 2s \text{ mol/L}$$

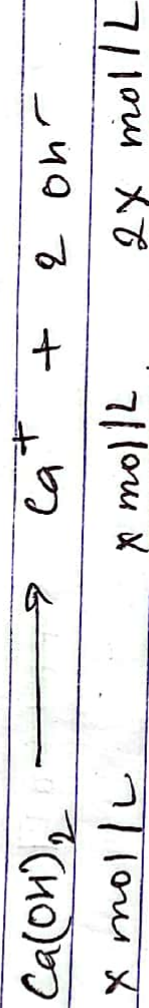
$$K_{sp} = [\text{Ca}^{++}]^2 [\text{OH}^-]^2 = s \times (2s)^2 = 4s^3$$

$$\therefore S = \left(\frac{K_{sp}}{4} \right)^{1/3}$$

$$S = \left(\frac{4.42 \times 10^{-5}}{4} \right)^{1/3}$$

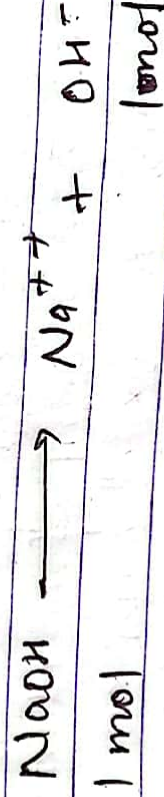
Solubility of $\text{Ca}(\text{OH})_2$ i.e. $S = 0.0223 \text{ mol/L}$.

Let solubility of $\text{Ca}(\text{OH})_2$ in $\text{NaOH} = x \text{ mol/L}$



Molarity of NaOH after mixing = $\frac{0.4}{2} \text{ M}$

[\therefore conc² becomes half when volume is made double].



$$[\text{OH}^-] = [\text{NaOH}] = 0.2 \text{ M}$$

\therefore Total concentration of $[\text{OH}^-] = [\text{OH}^-]$ from $\text{Ca}(\text{OH})_2 + [\text{OH}^-]$

~~from Na~~ from NaOH

$$\begin{aligned} K_{sp} \text{ of } \text{Ca}(\text{OH})_2 &= [\text{Ca}^{2+}] [\text{OH}^-]^2 \\ &= x \times (2x + 0.2)^2 \end{aligned}$$

x is very small i.e. $x < 0.2$. So, $2x + 0.2 \approx 0.2$

$$\therefore K_{sp} = x \times (0.2)^2$$

$$\therefore x = \frac{4.42 \times 10^{-5}}{0.04}$$

$$x = 0.0011 \text{ mol/L}$$

$$\text{we have, } M = \frac{g/L}{\text{mol. wt}}$$

$$\therefore g/L \text{ of } Ca(OH)_2 \text{ precipitate} = \left(\frac{s}{2} - x \right) \times \text{mol. wt}$$

$$= \left(\frac{0.0223}{2} - 0.0011 \right) \times 74$$

$$= 0.74 \text{ g}$$

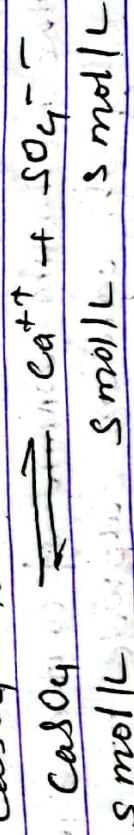
Q. What is the minimum volume of water required to dissolve 1 g of calcium sulphate at 298 K? (Given, solubility product K_{sp} for $CaSO_4 = 9.1 \times 10^{-6}$)

Given,

$$\text{Solubility of } CaSO_4 = 9.1 \times 10^{-6}$$

Minimum volume of water required = ?

$CaSO_4$ ionizes as,



CaSO_4 Molar Mass \rightarrow
 $40 + 32 + 64$
136

$$\therefore K_{sp} = [\text{Ca}^{++}] [\text{SO}_4^{--}]$$

$$= s \times s$$

$$s = \sqrt{K_{sp}}$$

$$= \sqrt{9.1 \times 10^{-6}}$$

$$= 3 \times 10^{-3} \text{ mol/L} \rightarrow \text{molarity}$$

$$\left[M = \frac{\text{g/L}}{\text{mol. wt.}} \right]$$

$$\text{g/L} = \text{Molarity} \times \text{Mol. wt.}$$

$$\therefore \text{Solubility of } \text{CaSO}_4 = 3 \times 10^{-3} \times 136 \text{ g/L.}$$

$$= 408 \times 10^{-3} \text{ g/L}$$

Since solute is very less. So vol. of solution is equal to vol. of solvent.

$\therefore 408 \times 10^{-3} \text{ g}$ of CaSO_4 is dissolved in 1000 ml of water.

$\therefore 1 \text{ g}$ of CaSO_4 is dissolved in $\frac{1000}{408 \times 10^{-3}}$ ml of water

$$= 2.45 \times 10^3 \text{ ml of water}$$

Hence, the minimum volume of water required to dissolve 1g of calcium sulphate at 298K is $2.45 \times 10^3 \text{ ml}$.