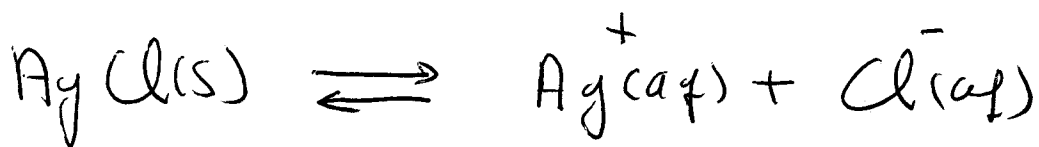


SOLUBILITY



$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.6 \times 10^{-10}$$

$$[\text{Ag}^+] = [\text{Cl}^-] = 1.3 \times 10^{-5}$$

- a) 500 mL OF A CaCl_2 SOL ($8.0 \times 10^{-6} \text{ M}$)
300 mL OF A AgNO_3 SOL ($4.0 \times 10^{-3} \text{ M}$)

$$[\text{Cl}^-] = 2 \times \frac{5}{8} (8.0 \times 10^{-6} \text{ M}) = 1.00 \times 10^{-5} \text{ M}$$

$$[\text{Ag}^+] = \frac{3}{8} (4.0 \times 10^{-3} \text{ M}) = 1.5 \times 10^{-3} \text{ M}$$

$$Q = 1.5 \times 10^{-8} \text{ M} > K_{sp}$$

\Rightarrow $[\text{Ag}^+]$ AND $[\text{Cl}^-] \downarrow \Rightarrow$ ppt $\text{AgCl}(s)$

COMMON-ION EFFECT

b) $\text{AgCl}(s)$ IS ADDED TO 1.00 L OF A 0.100 M SOL OF NaCl .

IF S IS THE SOLUBILITY OF $\text{AgCl}(s)$

$$[\text{Ag}^+] = S$$

$$[\text{Cl}^-] = 0.100 + S$$

$$K_{sp} = S(0.100 + S) \approx 0.100 S$$

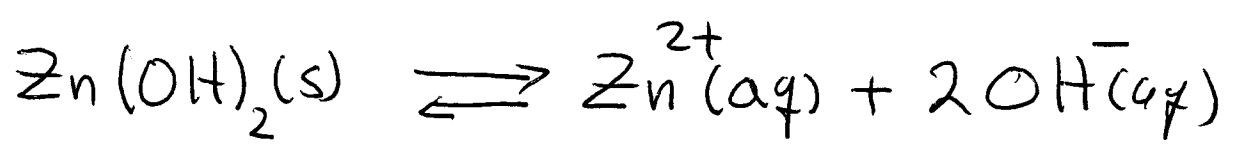
$$S = 1.6 \times 10^{-9} \text{ M} \ll 0.100$$

$$S \ll 1.3 \times 10^{-5} \text{ M (IN WATER)}$$

IN GENERAL IF $[\text{Cl}^-] \uparrow \Rightarrow [\text{Ag}^+] \downarrow$

EFFECT OF pH ON SOLUBILITY

CONSIDER $Zn(OH)_2(s)$



$$K_{sp} = [Zn^{2+}] [OH^-]^2$$

IF X IS THE SOLUBILITY OF $Zn(OH)_2(s)$,

WE HAVE

$$[Zn^{2+}] = X$$

$$[OH^-] = 2X$$

$$K_{sp} = 4X^3 \Rightarrow X = \sqrt[3]{\frac{K_{sp}}{4}}$$

$$X = [Zn^{2+}] = \sqrt[3]{\frac{4.5 \times 10^{-17}}{4}} = 2.2 \times 10^{-6} M$$

$$[OH^-] = 4.4 \times 10^{-6} M \Rightarrow pOH = 5.35$$

$pH = 8.65$

IF WE BUFFER THE SOLUTION TO CONTROL THE pH

$$K_{sp} = [Zn^{2+}] [OH^-]^2$$

BUT

$$K_w = [H_3O^+] [OH^-]$$

OR

$$[Zn^{2+}] = \frac{K_{sp}}{K_w^2} [H_3O^+]^2$$

$$= \frac{4.5 \times 10^{-17}}{1.0 \times 10^{-28}} [H_3O^+]^2$$

$$= 4.5 \times 10^{11} [H_3O^+]^2$$

$$\boxed{[Zn^{2+}] = 4.5 \times 10^{11} \times 10^{-2pH}}$$

$$\text{AT } pH = 6.00 \Rightarrow [Zn^{2+}] = 0.45 \text{ M } \uparrow$$