COMMON ION EFFECT IN SPARINGLY SOLUBLE SALT:

- The common ion effect decrease the solubility of the sparingly soluble salt.

- When a common ion is added, for eg: AgCl, According to le-Chatlier, equilibrium shifts towards backward direction and decrease the solubility of AgCl and favour the formation of more solid precipitate.

Que: Find out the solubility of $BaSO_4$ (Barium sulphate) ($K_{sp} = 10^{-4}$)
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*
(i) In Water
(ii) In 0.2M $Ba(NO_3)_2$ solution.

Sol^n: ① $BaSO_4(s) \xrightleftharpoons{H_2O} \underset{S}{Ba^{+2}_{(aq)}} + \underset{S}{SO_4^{-2}}(aq)$

$Ksp = S^2$
$10^{-4} = S^2$
$\therefore S = 10^{-2}$

② After adding $Ba(NO_3)_2$

$\underset{0.2M}{Ba(NO_3)_2} \longrightarrow \underset{0.2M}{Ba^{+2}} + \underset{0.2M}{2NO_3^-}$

Here, Ba^{+2} is common ion

So,

$$K_{sp} = [Ba^{+2}][SO_4^{-2}]$$

$$K_{sp} = (\overset{0}{\cancel{S}} + 0.2)\, S$$

$$K_{sp} = 0.2\, S$$

$$S = \frac{10^{-4}}{0.2}$$

$$\boxed{S = 5 \times 10^{-4}}$$

Que10 DPP07 ***
Solubility of CaF_2 ($K_{sp} = 3.4 \times 10^{-11}$) in 0.005M solution of BaF_2 is :

$$\underset{}{CaF_2(s)} \rightleftharpoons \underset{S}{Ca^{+2}(aq)} + \underset{2S}{2F^-(aq.)} \quad \text{— ①}$$

Add
$$\underset{0.005M}{BaF_2} \longrightarrow \underset{0.005\,M}{Ba^{+2}} + \underset{0.01\,M}{2F^-}$$

From reaction ①

$$K_{sp} = S\,(\overset{zero}{\cancel{2S}} + 0.01)^2$$

$$K_{sp} = 0.0001\, S$$

$$S = \frac{3.4 \times 10^{-11}}{0.001} = 3.4 \times 10^{-7} M$$

Que11 DPP07
K_{sp} of AgCl is 1×10^{-10}. Its solubility in 0.1 M KNO_3 will be

Strong electrolyte KNO_3 with AgCl, does not have common ion. so, solubility of AgCl does not effect after adding KNO_3.

Que12 Solubility of AgCl will be maximum in:
DPP:07

Salt	Common Ion
0.001 M $AgNO_3$	$Ag^+ = 0.01$ M
[✓ 0.01 M $CaCl_2$]	$Cl^- = 0.02$ M
0.01 M NaCl	$Cl^- = 0.01$ M

Concentration of common Ion ↑
↳ Solubility of sparingly soluble salt ↑

SIGNIFICANCE OF Ksp AND IONIC PRODUCT (I.P): or Q_{IP})

CASE:01 If $K_{sp} = Q_{IP}$: SATURATED SOLUTION

CASE:02 If $K_{sp} > Q_{sp}$:
In this case, solution is UNSATURATED, salt is more soluble in solvent and rxn moves from left to right (X precipitation

CASE:03 If $K_{sp} < Q_{sp}$
Precipitation should be occur or solubili of salt decrease or reaction moves backward direction till equilibrium.

Que 17 If Ksp (AgCl) is 10^{-10}, then which of the
DPP 07 solution are saturated with AgCl?

1) $[Ag^+] = 10^{-10}$, $[Cl^-] = 1M$
2) $(Ag^+) = 10^{-6} M$, $(Cl^-) = 10^{-5} M$
3) $(Ag^+) = 10^{-11}$, $(Cl^-) = 1M$
4) $(Ag^+) = 10^{-2} M$, $(Cl^-) = 10^{-8.5} M$

Que 16 Equal vol. of 2 sol^n are mixed. The one in which
DPP 07 $CaSO_4$ ($Ksp = 2.4 \times 10^{-5}$) is precipitated is

(1) $0.02M\ CaCl_2 + 0.004M\ Na_2SO_4$
(2) $0.2M\ CaCl_2 + 0.004M\ Na_2SO_4$
(3) $0.002M\ CaCl_2 + 0.04M\ Na_2SO_4$
(4) $0.2M\ CaCl_2 + 0.0004M\ Na_2SO_4$.

L.R. → Limiting Reagent.

$$\underset{0.02M}{CaCl_2} + \underset{0.004M}{\overset{(L.R.)}{Na_2SO_4}} \rightarrow \boxed{\underset{0.004M}{CaSO_4}} + \underset{0.004 \times 2}{2NaCl}$$

0.008 M

[Main Reaction]

$$CaSO_4(s) \rightarrow Ca^{+2}(aq) + SO_4^{-2}(aq)$$

$$\therefore Q_{IP} = (\cancel{s} + 0.004)(\cancel{s} + 0.004)$$

$$\boxed{Q_P = 16 \times 10^{-6} < Ksp_= 2.4 \times 10^{-5}}$$

ACID-BASE TITRATION

The titration (Volumetric Analysis) is carried out between acid and base is called ACID-BASE TITRATION.

INDICATOR

For Acid-Base titration, indicator is used to find out equivalence point.

INDICATOR	pH RANGE	COLOUR IN ACIDIC MEDIUM	COLOUR IN BASIC MEDIUM
* Methyl Orange	3.1 - 4.4	Pink	Yellow
Methyl Red	4.2 - 6.3	Red	Yellow
* Litmus	5.5 - 7.5	Red	Blue
Phenol Red	6.8 - 8.4	Yellow	Red
* Phenolphthalein	8.3 - 10	Colourless	Pink
Thymol Blue	1.2 - 2.8	Red	Yellow

- Indicators are weak organic acid or weak organic base. They show different colour in ionised and unionised form.

EQUIVALENCE POINT

The equivalence point is a fundamental concept in titration where the amount of acid and base that has been added are stoichiometrically till gram equivalent should be equal. It means acid and base are completely neutralised at equivalence point.

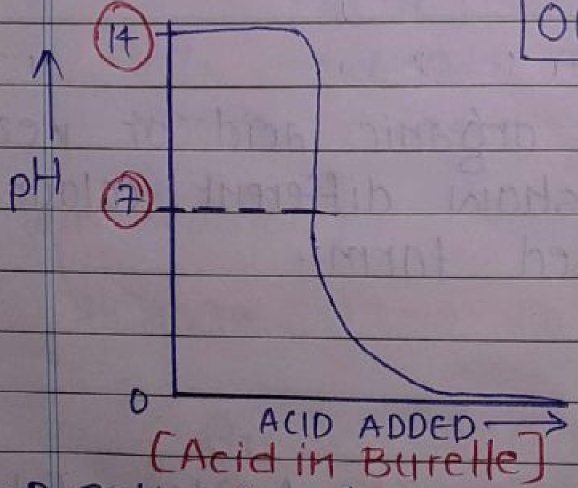
TITRATION	SASB	SAWB	SBWA	WAWB
Equivalence point pH	7	<7	>7	Can't determine

END POINT

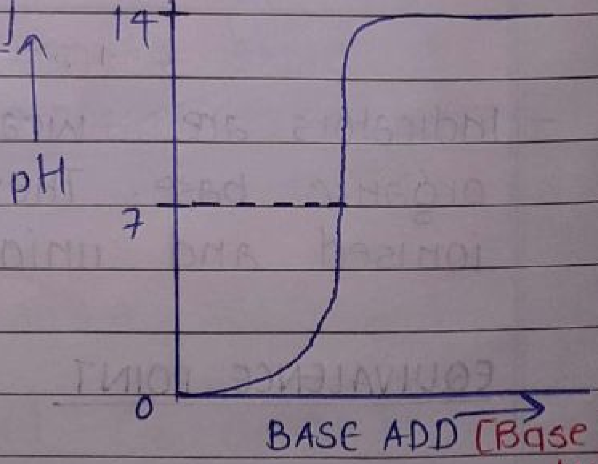
The point at which, indicator change the colour.

REPRESENTATION OF ACID-BASE TITRATION GRAPH

① SA and SB Eq: wt Point (SASB) → Neutral (=7)



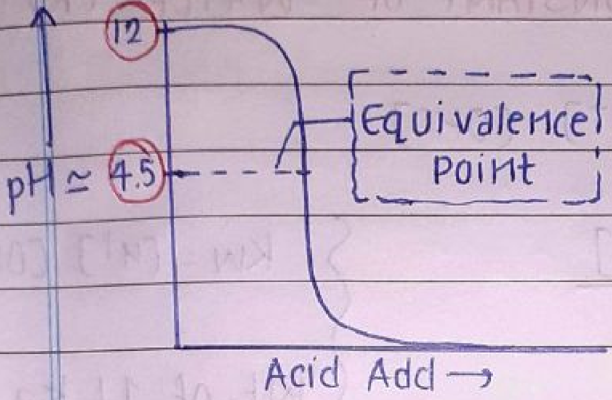
OR



[Base in burette]

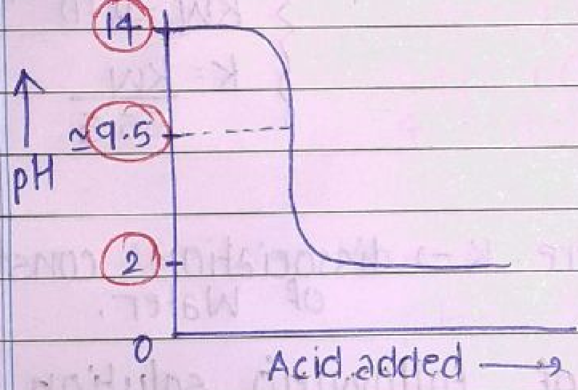
⇒ Both graph are correct for SA and SB.

② **SA and WB** Equivalent Point (SAWB = <7)
So ; $\boxed{4.5 < 7}$



[Acid in Burette]

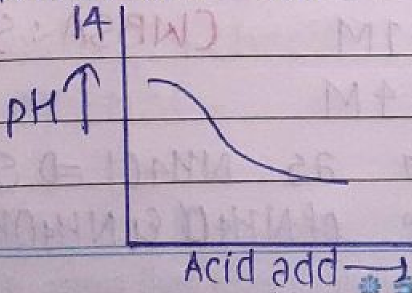
③ **SB and WA** Equivalent Point (SBWA = >7)
So ; $\boxed{9.5 > 7}$



[Acid in Burette]

④ **WA & WB** Equivalent Point (WAWB = Can't determine)

No sharp equivalence point observed during titration. So, generally ~~they~~ that type of titration does not carried out in titration.



dissociation constant of Water.

Que: Find out the pH of following solution after mixing NH_4OH [0.5M] and HCl [0.1M].
5*
Given $K_b = 10^{-5}$

Solution:

(L.R. → Limiting Reagent)

$$NH_4OH + \overset{\downarrow}{HCl} \longrightarrow NH_4Cl + H_2O \quad \text{(Before reaction)}$$

$$0.5M \qquad 0.1M \qquad\qquad 0.1M$$

After reaction, $HCl = 0$

$NH_4Cl = 0.1M$ [WBSA : Salt]

$NH_4OH = 0.4M$

Yes, it is Buffer solution as NH_4Cl ⇒ SA and NH_4OH ⇒ WB. If the value of NH_4Cl & NH_4OH is zero then it is not Buffer Solution.

Using formula,

$$pOH = pK_b + \log \frac{[Salt]}{\cancel{[Acid]}\ [WB]}$$

$$\begin{array}{r} 4\,9\,9\,10 \\ 5.0000 \\ 0.6020 \\ \hline 4.3980 \end{array}$$

$$\therefore pOH = 5 + \log \frac{[0.1]}{[0.4]}$$

$$\Rightarrow pOH = 5 + [\log 1 - \log 4]$$
$$= 5 + [-\log 4]$$
$$= 5 + [-0.6020]$$
$$= 4.3980$$

$\{ \log 1 = 0$

$\log 4 = 0.6020$

$$\begin{array}{r} 13 \\ 3\,9\,9\,10 \\ 14.0000 \\ 4.3980 \\ \hline 9.6020 \end{array}$$

As we know,

$$pH + pOH = 14$$

So,

$$pH = 14 - 4.3980$$
$$= 9.6020$$

$\{ pOH = 4.3980$

Final answer; $\boxed{pH = 9.6020}$

Que: 0.1M NH_4OH mixed with 0.5M HCL. Then justify / check whether it is Buffer solution or not.

6*

Solution: Next Page →

limiting reagent.

Solution: (L.R.)

$$\underset{0.1M}{NH_4OH} + \underset{0.5M}{HCl} \longrightarrow NH_4Cl + H_2O$$ [Before Reaction]

After reaction, $NH_4OH = 0$
$HCl = 0.4M$
$NH_4Cl = 0.1M$

Hence, it is not a buffer solution because concentration of weak base (NH_4OH) is zero after completion of reaction.