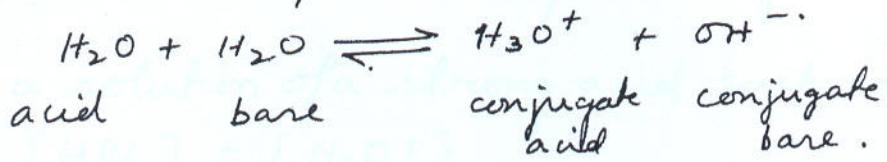


(10)

- ~~Ques.~~ 1. Concepts and definitions used in ionic equilibria
- Pure water undergoes self ionization



- $[\text{H}_3\text{O}^+] \times [\text{OH}^-] = K_w$ is the ionic product of water.
- $K_w = 1.0 \times 10^{-14}$ at 298 K.
At higher temperatures the value increases.

- The reaction opposite to self ionization is called neutralization. Pure water is neutral.
- A neutral solution is where $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 10^{-7}$
- An acidic solution is where $[\text{H}_3\text{O}^+] > [\text{OH}^-]$
- A basic or alkaline solution is where $[\text{H}_3\text{O}^+] < [\text{OH}^-]$
- Strong electrolytes ionize to a large extent in water.
- Weak electrolytes ionize weakly in water.
- Non electrolytes do not ionize in water.

- According to:
- The general Lowry-Bronsted theory, when a molecule A transfers a proton to another molecule B, A is called an acid and B is called a bare. The products are conjugate base of A and conjugate acid of B respectively.

- Water can both accept and donate protons. In aqueous solutions water is taken as the standard.
- Substances which donate protons to water are called acids and substances which accept protons from water are called bases.

- Acids which are strong electrolytes are strong acids and bases which are strong electrolytes are strong bases.

In a solution of a strong acid such as HCl.

$$\underline{[HCl] = [H_3O^+]}$$

In a solution of a strong base such as NaOH.

$$\underline{[NaOH] = [OH^-]}$$

(HCl is a Brønsted acid but NaOH is not a Brønsted base. The OH^- produced from NaOH is the actual Brønsted base)

- The p_H of a solution is given by the expression

$$\boxed{p_H = -\log [H_3O^+]} \quad (\text{The actual definition is } p_H = -\log a_{H_3O^+})$$

where the activity $a_{H_3O^+} \approx [H_3O^+]$ for dilute solutions.

The expression $-\log [OH^-]$ is called p_{OH}

and the expression $-\log K_w$ is called p_{Kw} .

- $[H_3O^+] [OH^-] = K_w = 1 \times 10^{-14}$.

Hence.

$$\boxed{p_H + p_{OH} = p_{Kw} = 14}$$

- when solution is neutral.

$$\boxed{p_H = p_{OH} = 7}$$

- when solution is acidic.

$$p_H < p_{OH}, \boxed{p_H < 7} \text{ and } p_{OH} > 7.$$

- when solution is basic.

$$p_H > p_{OH}, \boxed{p_H > 7} \text{ and } p_{OH} < 7.$$

- In normal cases the p_H of a solution varies between 0 and 14. However, concentrated solutions of strong acids and bases may have p_H values less than zero and greater than 14 respectively. At these concentrations $a_{H_3O^+}$ is usually different from the $[H_3O^+]$ value.
- When electrolytes are dissolved in water, the ions produced, often react with water.
acids donate protons to water. $HA + H_2O \rightleftharpoons A^- + H_3O^+$
bases accept protons from water. $B + H_2O \rightleftharpoons BH^+ + OH^-$
- Conjugate bases and acids behave ~~as~~ as their name suggests
 A^- acts as a base $A^- + H_2O \rightleftharpoons HA + OH^-$
 BH^+ acts as an acid $BH^+ + H_2O \rightleftharpoons B + H_3O^+$.
- The strengths of acids and bases are measured in terms of the equilibrium constants of their dissociations, K_a and K_b respectively.

For the acid HA :

$$K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

For the base B :

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

For the conjugate base A^- of the acid HA :

$$K_b' = \frac{[HA][OH^-]}{[A^-]}$$

For the conjugate acid BH^+ of the base B :

$$K_a' = \frac{[B][H_3O^+]}{[BH^+]}$$

K_a of HA and K_b' of A^- are related.

$$K_a K_b' = K_w \text{ or. } K_b' = \frac{K_w}{K_a}$$

Similarly, K_b of B and K_a' of BH^+ are related

$$K_b K_a' = K_w \text{ or. } K_a' = \frac{K_w}{K_b}$$

Thus if K_a is high K_b is low
and if K_b is high K_a' is low.

- Acids and bases are classified as

strong $\Rightarrow K$ is very high and > 1

moderately weak $\Rightarrow K$ is low but $> 10^{-7}$.

very weak $\Rightarrow K$ is very low and $< 10^{-7}$.
but $> 10^{-14}$.

Ineffective $\Rightarrow K$ is insignificant and $\leq 10^{-14}$.

- Acids and bases react with each other to produce salts whose ions are

conjugate base of the acid
and conjugate acid of the base.

when the acid or the base are strong, moderately weak, very weak or ineffective, their respective conjugates are weak, moderately weak, ineffective, very weak, weak or strong respectively

- Salt of a strong acid with a weak base \equiv ineffective conjugate base + weak conjugate acid \equiv a weak acid due to hydrolysis. (example NH_4Cl)

• Salt of a strong base with a weak acid \equiv ineffective conjugate acid + weak conjugate base \equiv a weak base due to hydrolysis (example $\text{CH}_3\text{CO}_2\text{Na}$)

• Salt of a weak base with a weak acid \equiv weak conjugate acid + weak conjugate base \equiv weak acid or weak base depending on whether the acid or the base is stronger (example $\text{CH}_3\text{CO}_2\text{NH}_4$)

of the two.

- salt of a strong acid with a strong base. \equiv ineffective conjugate base + ineffective conjugate acid \equiv neutral solution. (example NaCl).
- The fraction of an acid or base that gets dissociated is called the degree of dissociation ' α '.
The dissociation of conjugate acids or bases ~~are~~ is also called hydrolysis. The fraction of molecules (or ions) that get hydrolysed are accordingly called degree of hydrolysis ' h '.

2. Determination of $[H_3O^+]$ in solutions of acids, bases or salts.

- For strong acids and bases $\alpha \approx 1$. For weak acids and bases (or conjugate acids and bases) $\alpha < 1$.
The value of ' α ' is dependent on the constant of dissociation as well as the concentration of the substance being dissociated as.

$$K_a \text{ (or } K_a') = \frac{c\alpha^2}{1-\alpha} \approx c\alpha^2 \quad (\text{when } \alpha \ll 1).$$

$$K_b \text{ (or } K_b') = \frac{c\alpha^2}{1-\alpha} \approx c\alpha^2 \quad (\text{when } \alpha \ll 1).$$

This relation is also called "Ostwald's dilution law".

The general expression is

$$\text{or. } \alpha = \sqrt{\frac{K}{c}}$$

(where $K = K_a, K_a', K_b, K_b'$ depending on the identity of the substance).

- For solutions of acids $[H_3O^+] = c\alpha$

For strong acids $[H_3O^+] \approx c$ (since $\alpha \approx 1$). (eg HCl)

For weak acids $[H_3O^+] = c\alpha$ (eg. CH_3CO_2H).

$$= c\sqrt{\frac{K_a}{c}} \quad (\text{when } \alpha \ll 1)$$

$$\text{or. } [H_3O^+] = \sqrt{cK_a}.$$

$$\text{and } \left(p_H = \frac{1}{2} p_{K_a} - \frac{1}{2} \log c \right)$$

- For solutions of bases: $[\text{OH}^-] \approx c\alpha$.

For strong bases $[\text{OH}^-] = c$ (since $\alpha \approx 1$) (eq. NaOH)

For weak acids $[\text{OH}^-] = c\alpha$ (eq. NH_3).

$$= c \sqrt{\frac{K_b}{c}} \quad (\text{when } \alpha \ll 1)$$

or $[\text{OH}^-] = \sqrt{c K_b}$

and. $\text{p}_{\text{OH}} = \frac{1}{2} \text{p}_{K_b} - \frac{1}{2} \log c$

Also. $\text{p}_{\text{H}} = \text{p}_{K_w} - \text{p}_{\text{OH}} = \text{p}_{K_w} - \frac{1}{2} \text{p}_{K_b} + \frac{1}{2} \log c$

For solutions of salts

- Salt of strong acid with weak base (eq. $\text{NH}_3 \text{Cl}$)

$$[\text{H}_3\text{O}^+] = \sqrt{c K_a} = \sqrt{c \frac{K_w}{K_b}}$$

$$\text{p}_{\text{H}} = \frac{1}{2} \text{p}_{K_a} - \frac{1}{2} \log c = \frac{1}{2} \text{p}_{K_w} + \frac{1}{2} \text{p}_{K_b} - \frac{1}{2} \log c$$

Also degree of hydrolysis $h = \sqrt{\frac{K_a}{c}} = \sqrt{\frac{K_w}{K_b c}}$

- Salt of strong base with weak acid (eq. $\text{CH}_3\text{CO}_2\text{Na}$)

$$[\text{OH}^-] = \sqrt{c K_b} = \sqrt{c \frac{K_w}{K_a}}$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \sqrt{\frac{K_w K_a}{c}}$$

$$\text{p}_{\text{H}} = \frac{1}{2} \text{p}_{K_w} + \frac{1}{2} \text{p}_{K_a} + \frac{1}{2} \log c$$

Also degree of hydrolysis $h = \sqrt{\frac{K_b}{c}} = \sqrt{\frac{K_w}{K_a c}}$

• For solutions of salts of weak acids with weak bases

(e.g. $\text{CH}_3\text{CO}_2\text{NH}_4^+$)

For the salt of weak acid HA with weak base B the salt has the formula $(\text{BH}^+)\text{A}^-$ ionizing as BH^+ and A^- . Unlike in the cases discussed above BH^+ and A^- both hydrolyse and the H_3O^+ and OH^- get neutralized. This means that $[\text{H}_3\text{O}^+] \neq c_h = \sqrt{cK_a}$ for BH^+ and $[\text{OH}^-] \neq c_h = \sqrt{cK_b}$ for A^- .

In such cases, HA , B , A^- , ~~BH^+~~ and OH^- are all present related by the charge balance relation.

$$[\text{H}_3\text{O}^+] + [\text{BH}^+] = [\text{OH}^-] + [\text{A}^-].$$

and the mass balance relation

$$[\text{BH}^+] + [\text{B}] = [\text{A}^-] + [\text{HA}].$$

when the ~~salt~~ solution is not very dilute.

$[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ are $\ll [\text{BH}^+]$ and $[\text{A}^-]$.
and $[\text{BH}^+] = [\text{A}^-]$, $[\text{B}] = [\text{HA}]$.

In such cases both $[\text{H}_3\text{O}^+]$ as well as p_h are independent of concentration of the salt.

$$[\text{H}_3\text{O}^+] = \sqrt{\frac{K_w K_a}{K_b}}$$

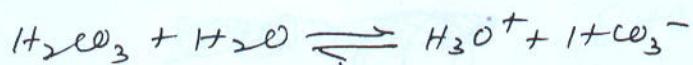
$$\text{and } p_h = \frac{1}{2} p_{K_w} + \frac{1}{2} p_{K_a} - \frac{1}{2} p_{K_b}$$

$$\text{Also } \frac{h}{1-h} \approx h = \sqrt{\frac{K_w}{K_a K_b}} \quad (\text{when } h \ll 1).$$

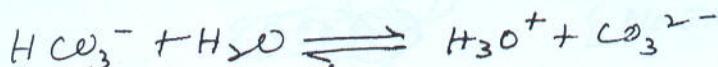
- For solutions of acid salts of weak polybasic acids with strong bases. (e.g. $\text{Na}^+ \text{CO}_3^-$)

This behaves as a salt of a weak acid (H_2CO_3) and a weak base (CO_3^{2-} which is a conjugate base).

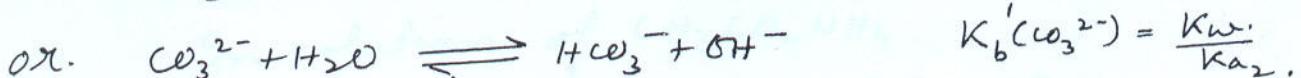
~~H_2CO_3~~ H_2CO_3 ionizes in two steps.



$$K_{a_1} = K_a(\text{H}_2\text{CO}_3)$$



$$K_{a_2}$$



As a corollary of the previous case we have.

$$[\text{H}_3\text{O}^+] = \sqrt{\frac{K_w K_a(\text{H}_2\text{CO}_3)}{K_b'(\text{CO}_3^{2-})}} = \sqrt{\frac{K_w K_{a_1}}{K_w/K_{a_2}}} = \sqrt{K_{a_1} K_{a_2}}$$

or
$$[\text{H}_3\text{O}^+] = \sqrt{K_{a_1} K_{a_2}}$$

$$pH = \frac{1}{2} pK_{a_1} + \frac{1}{2} pK_{a_2}$$

and $h = \sqrt{\frac{K_w}{K_a(\text{H}_2\text{CO}_3) K_b'(\text{CO}_3^{2-})}} = \sqrt{\frac{K_w}{K_{a_1} K_w/K_{a_2}}}$

or
$$h = \sqrt{\frac{K_{a_2}}{K_{a_1}}}$$

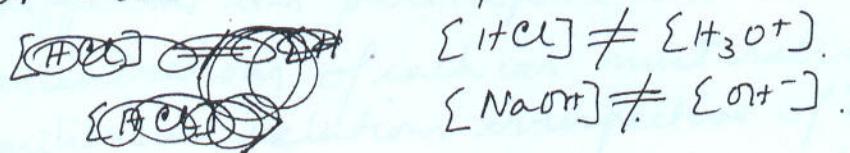
Here too, both $[\text{H}_3\text{O}^+]$ and 'h' are independent of concentration.

For solving problems.

- Points to remember while solving problems.

- In many of the formulas derived, α (or h) is considered to be $\ll 1$. This approximation fails when K_a or K_b is not very small or the solution is very dilute. In such cases the simple formulas cannot be used.

- When the solution is very dilute, i.e. the $\{\text{H}_3\text{O}^+\}$ and $\{\text{OH}^-\}$ are not much less than (i.e. comparable to) the concentrations of the solute then ~~approximations~~. They have to be taken into consideration for all calculations and most of the formula's ~~do not~~ are ~~not~~ not valid. For example.



In solutions of $\text{CH}_3\text{CO}_2\text{NH}_4$.

$$\{\text{CH}_3\text{CO}_2^-\} \neq \{\text{NH}_4^+\}$$

etc.

Hence in all problems, the validity of assumptions must be checked. The approximations must be correct within 10%.

3. ~~Remember~~ Trying to remember so many formula's may create problems. Hence it is advisable to remember the basic formulas and make necessary substitutions later as shown below.

$\text{CH}_3\text{CO}_2\text{H}$: $\boxed{\{\text{H}_3\text{O}^+\} = \sqrt{cK_a}}$; $\boxed{\alpha = \sqrt{\frac{K_a}{c}}}$ (Calculate $\{\text{H}_3\text{O}^+\}$ from $\{\text{OH}^-\}$)

NH_3 : $\boxed{\{\text{OH}^-\} = \sqrt{cK_b}}$; $\boxed{\alpha = \sqrt{\frac{K_b}{c}}}$ (Calculate $\{\text{H}_3\text{O}^+\}$ from $\{\text{OH}^-\}$)

NH_4^+ (as in NH_4Cl): $\boxed{\{\text{H}_3\text{O}^+\} = \sqrt{cK_a'}}$ or. $\sqrt{cK_h}$. ($K_h = K_a'$ of NH_4^+ .
 $= \frac{K_w}{K_b}$ of NH_3 .)

$$h(\text{or } \alpha) = \sqrt{\frac{K_a'}{c}} \text{ or. } \sqrt{\frac{K_h}{c}}$$

CH_3CO_2^- (as in $\text{CH}_3\text{CO}_2\text{Na}$): $\boxed{\{\text{OH}^-\} = \sqrt{cK_b'}}$ or. $\sqrt{cK_h}$. ($K_h = K_b'$ of CH_3CO_2^- .
 $= \frac{K_w}{K_a}$ of $\text{CH}_3\text{CO}_2\text{H}$.)

Also calculate $\{\text{H}_3\text{O}^+\}$ from $\{\text{OH}^-\}$.

$\text{CH}_3\text{CO}_2\text{NH}_4$: $\boxed{\{\text{H}_3\text{O}^+\} = \sqrt{\frac{K_w K_a}{K_b'}}$; $\boxed{h(\text{or } \alpha) = \sqrt{\frac{K_w}{K_a K_b'}}$

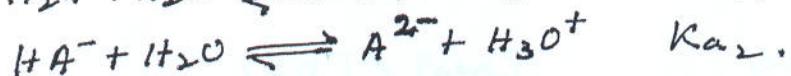
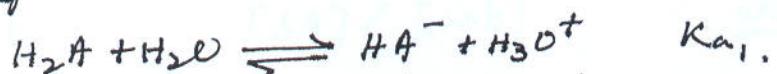
NaHCO_3 : $\boxed{\{\text{H}_3\text{O}^+\} = \sqrt{K_{a_1} K_{a_2}}}$; $\boxed{h(\text{or } \alpha) = \sqrt{\frac{K_{a_2}}{K_{a_1}}}}$

3. Determination of $[H_3O^+]$ - 10 - (19)

For a combination of solutes.

- When two different solutions are mixed ~~also~~ adjust the concentration of each solute for the final volume.
- Consider the different reactions and equilibria the different ions participate in. Each ion must
- The concentrations of each ion must satisfy all the equilibrium relations irrespective of their origin. If the same ion is formed from two different solutes then we have the common ion effect. If one of the solutes produce a large excess of the "common ion" then the dissociation of the other solute is suppressed.

Corollary. When a dibasic acid ionizes.



Then (1) $K_{a_1} = \frac{[HA^-][H_3O^+]}{[H_2A]}$ determines the $[H_3O^+]$

(2) $[HA^-] = [H_3O^+] = \sqrt{cK_{a_1}}$ (provided K_{a_1} is not very large).

and (3) $[A^{2-}] = K_{a_2}$ (provided $K_{a_2} \gg K_{a_1}$).

- If both acids as well as bases are present in the solution, consider the reaction $H_3O^+ + OH^- \rightleftharpoons 2H_2O$ in order to evaluate the final $[H_3O^+]$. For doing this proper stoichiometric calculations are necessary.

- Individual components may be categorized as. SA (strong acid eg HCl), SB (strong base eg NaOH). wa (weak acid eg CH_3CO_2H or NH_4Cl), wb (weak base eg NH_3 or CH_3CO_2Na) or. wawb (combination of

a weak acid and a weak base which behaves as a buffer (e.g. $\text{CH}_3\text{CO}_2\text{NH}_3$ or NaHCO_3). Then the combinations also fall into any one of the categories depending on the relative concentrations of the combinations.

<u>Combination</u>	<u>Relative concentration</u>	<u>Resultant category</u>
$\text{SA} + \text{wa}$	—	SA
$\text{SB} + \text{wb}$	—	SB
$\text{SA} + \text{SB}$	$[\text{SA}] > [\text{SB}]$	SA
$\text{SA} + \text{SB}$	$[\text{SA}] < [\text{SB}]$	SB
$\text{wa} + \text{wb}$	—	<u>wawb</u> . (buffer)
$\text{SA} + \text{wb}$	$[\text{SA}] > [\text{wb}]$	SA ($\text{SA} + \text{wa}$)
$\text{SA} + \text{wb}$	$[\text{SA}] = [\text{wb}]$	wa
$\text{SA} + \text{wb}$	$[\text{SA}] < [\text{wb}]$	<u>wawb</u> . (buffer)
$\text{SB} + \text{wa}$	$[\text{SB}] > [\text{wa}]$	SB ($\text{SB} + \text{wb}$)
$\text{SB} + \text{wa}$	$[\text{SB}] = [\text{wa}]$	wb
$\text{SB} + \text{wa}$	$[\text{SB}] < [\text{wa}]$	<u>wawb</u> . (buffer)
$\text{SA} + \text{wawb}$	$[\text{SA}] < [\text{wawb}]$	<u>wawb</u> . (buffer)
$\text{SA} + \text{wawb}$	$[\text{SA}] = [\text{wb}]$	wa^*
$\text{SA} + \text{wawb}$	$[\text{SA}] > [\text{wawb}]$	SA ($\text{SA} + \text{wa}$)
$\text{SB} + \text{wawb}$	$[\text{SB}] < [\text{wawb}]$	<u>wawb</u> . (buffer)
$\text{SB} + \text{wawb}$	$[\text{SB}] = [\text{wawb}]$	wb^*
$\text{SB} + \text{wawb}$	$[\text{SB}] > [\text{wawb}]$	SB ($\text{SB} + \text{wb}$)

* This relation may not hold when the 'wawb' compound is a salt of a polyprotic acid.

(2)

Buffer solutions

4. Acid base titrations

- The concentrations of acids or bases in an "unknown" solution (concentration is unknown) may be experimentally be determined by titrating it with a solution of bases or acids respectively of known concentrations. When $\text{molar eq acid} = \text{molar eq base}$ then the solution is neutral if K_a & K_b are equal.
- When $\text{molar eq acid} = \text{molar eq base}$ it is called the equivalence point of a titration.
- Obviously ~~the~~. The solution is neutral at the equivalence point when both ~~are~~ the acid and the base are ~~weak~~ strong. The pH of the solution changes sharply from 4 to 10 at the equivalence point depending on whether the acid or the base is in slight excess.
- If the acid is strong but the base is weak, then the pH at the equivalence point is less than 7. If the base is strong but the acid is weak, then the pH at the equivalence point is greater than 7. Accordingly the sharp change in pH takes place in the acidic or the basic region.
- The end point of a titration is observed by using an indicator which is a weak acid having different colours as HA and A^- respectively. The indicator changes colour within the range $pK_a(\text{Indicator}) \pm 1$.

$$pH = pK_a(\text{Indicator}) \pm 1.$$

- The range of the indicator must fall within the range of sharp p_H change in the titration for the accurate determination of end points.

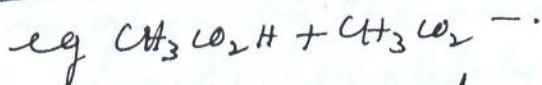
5. Buffer solutions

- Buffers are solutions of the weak type, i.e., those having a combination of a weak acid or a weak base.

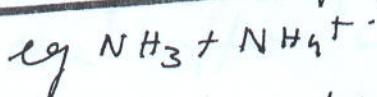
- The p_H of a buffer does not change too much with dilution or addition of small amounts of acids or bases.

- Buffers are of four types.

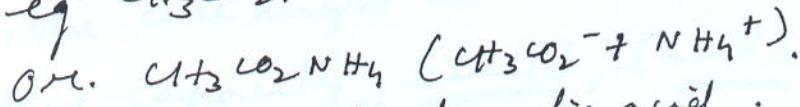
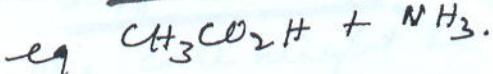
- (1) weak acid + conjugate base.



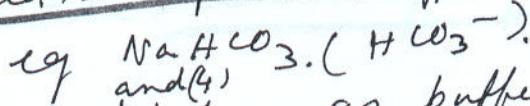
- (2) weak base + conjugate acid.



- (3) ~~weak acid + weak base~~.



- (4) acid salt of weak polyprotic acid.



- Buffer types (3) behave as buffer type (1) if a strong acid is added and as buffer type (2) if a strong base is added.

- In both buffer types (1) ~~and~~ (weak acid and its salt) and (2) (weak base and its salt) the acids, bases and their respective conjugates present as salts do not dissociate or hydrolyze significantly.

and the $[H_3O^+]$ or $[OH^-]$ can be calculated as.

• For buffer type(1).

$$[H_3O^+] = K_a \frac{[\text{acid}]}{[\text{salt}]}$$

$$\text{or } p_H = p_{K_a} + \log \frac{[\text{salt}]}{[\text{acid}]}$$

• For buffer type (2)

$$[OH^-] = K_b \frac{[\text{base}]}{[\text{salt}]}$$

$$\text{or } p_H = p_{K_b} + \log \frac{[\text{salt}]}{[\text{base}]}$$

• Buffer range is given by

$$p_H = p_{K_a} \pm 1$$

for type(1)

$$p_{OH} = p_{K_b} \pm 1$$

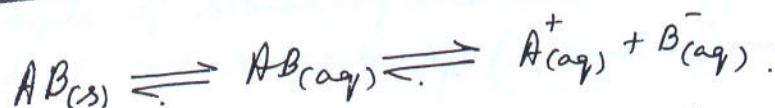
for type(2).

• Buffer capacity is given by the amount of acid or base which reacts completely with any one of the two components of the buffer present in one litre of the buffer solution.

- For maximum buffer range $[\text{salt}] = [\text{acid}]$.
- or $[\text{base}]$ as the case might be.
- For high buffer capacity $[\text{salt}]$, $[\text{acid}]$.
- or $[\text{base}]$ must be reasonably high.

6. Solubility and solubility product.

- When ionic solids are added to water, the ion from the dissolved salts enter into a heterogeneous equilibrium with the undissolved salts in a saturated solution.



- The maximum equilibrium value of the ionic product $[A^{+}_{(aq)}][B^{-}_{(aq)}]$ is called the solvability product K_{sp} of AB .

- For sparingly soluble salts condition for precipitation when a precipitating reagent is added is $\text{ionic product} > \text{solvability product}$.

- Molar solubility determines the ~~solute~~ value of K_{sp} of sparingly soluble salts.
- For salts of type AB :
 $[A^+] = [B^-] = s \Rightarrow K_{sp} = [A^+][B^-] = s^2$.
- Similarly for salts of type AB_2 :

$$[A^{2+}] = s, [B^-] = 2s.$$

$$\Rightarrow K_{sp} = [A^+][B^-]^2 = s(2s)^2 = 4s^3.$$

This relation also holds for salts of type A_2B .

For salts of type: AB_3 ; $K_{sp} = 27s^3$

For salts of type. A_2B_3 or A_3B_2 , $K_{sp} = 108s^5$.

- The solvability of a sparingly soluble salt is reduced due to common ion effect if the solution already has dissolved salts with an ion common with the sparingly soluble salt. For example AgCl is more soluble in water than in a solution of KCl .

- When two ~~soluble~~ sparingly soluble salts ~~are dissolved~~ ~~in~~ having a common ion are simultaneously dissolved then the ratio of the solubilities are determined by the ratio of K_{sp} values. The relationship depends on the formulae of the salts.

For example. if AB has solubility s_1 and solubility product K_{sp_1}

For example if solubility and solubility products of AB and AC are respectively. (s_1 and s_2) and (K_{sp_1} and K_{sp_2})

$$\text{Then. } \frac{K_{sp_1}}{K_{sp_2}} = \frac{[A^+][B^-]}{[A^+][C^-]} = \frac{s_1}{s_2}.$$

But for AB_2 and AC_2

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$$\frac{K_{sp,1}}{K_{sp,2}} = \frac{[A^{2+}] [B^-]^2}{[A^{2+}] [C^-]^2} = \frac{(2s_1)^2}{(2s_2)^2} = \frac{s_1^2}{s_2^2}$$

etc.

- Salts of weak acids have anions which are weak conjugate bases which hydrolyse in water.
The extent of hydrolysis may be controlled by externally adjusting the p_H of the solution. A low p_H ensures greater hydrolysis which implies greater solubility of the salts.
- The cations of many salts can be made to form complexes with reagents such as NH_3 , CN^- , $S_2O_3^{2-}$ etc. The solubility of such salts are enhanced by the addition of proper complexing agents.
- Saturating a solution with H_2S leads to the formation of S^{2-} ions where the p_H of the solution determines the $[S^{2-}]$ concentration. Selective precipitation of sparingly soluble metal sulphides is possible by controlling the p_H of the solution.

If K_1 and K_2 are the first and second dissociation constants of aqueous H_2S and ' s ' be its concentration in a saturated solution

$$\text{Then. } \frac{[S^{2-}] [H_3O^+]^2}{[H_2S]} = K_1 K_2.$$

or.
$$[S^{2-}] = \frac{K_1 K_2 (s)}{[H_3O^+]^2}.$$