Chemistry
Chapter: 8
Acids, Bases and Salts

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→ **Introduction:-**

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i) Acids

1) Acids produce H^+ ions when dissolved in H_2O.

$$\boxed{HCl \xrightleftharpoons{H_2O} H^+ + Cl^-}$$

2) Acids are sour in taste.

3) They are classified into a) mineral acids
b) organic acids

Organic acids are weaker and are found in vegetables, fruits etc.

ORGANIC ACIDS	WHERE IT IS FOUND
i. lactic acid	sour milk
ii. citric acid	citrus fruits
iii. formic acid	insect bites
iv. tartaric acid	grape juice
v. maleic acid	apples and pears

ii) Base

1) Base produces OH^- ions when dissolved in water.

$$\boxed{NaOH \xrightleftharpoons{H_2O} Na^+ + OH^-}$$

2) It has a bitter taste.

3) It has a slippery feel.

4) When a base is dissolved in water, an alkali is produced.

$$\boxed{CaO + H_2O \rightarrow Ca(OH)_2}$$

Here CaO is base and $Ca(OH)_2$ is an alkali.

5) Alkalies are used in making soap and detergents.

iii) Salt

1) When the neutralization takes place between an acid and base, salt is produced.

$$\underset{(acid)}{HCl} + \underset{(base)}{NaOH} \rightarrow \underset{(salt)}{NaCl} + \underset{(water)}{H_2O}$$

2) A salt may be neutral, acidic or basic.

→ Acidic, Basic and Amphoteric Substances :-

Acidic Substance:

1) A substance which gives acidic solution when dissolved in water is acidic e.g.

2) "oxides of non-metals" e.g., CO_2, SO_2 and N_2O_5 produce H_2CO_3, H_2SO_3 and HNO_3 when dissolved in water.

3) $CO_2 + H_2O \rightarrow H_2CO_3$

$SO_2 + H_2O \rightarrow H_2SO_3$

$N_2O_5 + H_2O \rightarrow 2HNO_3$

Basic Substance:

1) A substance which gives basic solution when dissolved in water is basic e.g.

2) "oxides of metals" like CaO and Na_2O when dissolved in water produce

NaOH and $Ca(OH)_2$.

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3) $Na_2O + H_2O \rightarrow 2NaOH$

$CaO + H_2O \rightarrow Ca(OH)_2$

Amphoteric Substance:

1) Some substances when dissolved in water produce such solutions which are of amphoteric nature i.e behave as acid as well as base. Such substance is called amphoteric substance. 2) "Less electropositive elements" are amphoteric substance.

3) Al_2O_3

$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$ (as base)

$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$ (as acid)

4) Cr_2O_3

$Cr_2O_3 + 6HNO_3 \rightarrow 2Cr(NO_3)_3 + 3H_2O$ (as base)

$Cr_2O_3 + 6KOH \rightarrow 2K_3CrO_3 + 3H_2O$ (as acid)

- Acidic Salts:-

1) Salts which produce acid when dissolved in water are of acidic nature.

2) e.g. NH_4Cl

$NH_4Cl + H_2O \rightarrow NH_4OH + HCl$.

- Basic Salts:-

1) Salts which produce bases when dissolved in water are of basic nature.

2) e.g. K_2CO_3

$K_2CO_3 + H_2O \rightarrow KOH + H_2CO_3$

→ Bronsted-Lowery concepts for acids and bases :-

Bronsted-Lowery Acid :-

Bronsted Acid is anything that donates proton or hydrogen ion (H^+).

Example:-

i. $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

In this reaction, the HCl donate H^+ to water and becomes Cl^- so it is an acid here while water would be base.

Bronsted-Lowery Base :-

Bronsted Base is anything that accepts proton or hydrogen ion (H^+).

Example:-

i. $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$

In this reaction, water accepts H^+ and becomes H_3O^+, so it acts as a base.

ii. $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

In this reaction, water donates H^+ which is accepted by NH_3 to become NH_4^+, so here NH_3 acts as a base.

Note:-

Amphoteric Substance:-

There are substances like H_2O and CH_3COOH which act as both acid and base. Such substances are called as amphoteric.

H_2O

$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$ (as base)

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$ (as acid)

CH_3COOH

$CH_3COOH + H_2O \rightarrow CH_3COO^- + H_3O^+$ (as acid)

$CH_3COOH + HCl \rightarrow CH_3COO^+H_2 + Cl^-$ (as base)

→ **Salt, Conjugated Acids and Conjugate Bases:-**

Salt:-

The substance obtained as a result of neutralization of acids and bases are called salts.

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + H_2O$$

Types

i Neutral Salt

It is obtained by neutralization of strong acid and strong base.

$$\underset{(\text{strong})}{HCl} + \underset{(\text{strong})}{NaOH} \rightarrow \underset{(\text{neutral})}{NaCl} + H_2O$$

ii Acid Salt

It is obtained by: the neutralisation of a strong acid with a weak base.

$$\underset{\text{strong acid}}{2HCl} + \underset{\text{weak base}}{Ca(OH)_2} \rightarrow \underset{\text{acidic salt}}{CaCl_2} + H_2O$$

iii Basic Salt

It is obtained by the reaction of a strong base with a weak acid.

$$\underset{\text{strong base}}{2NaOH} + \underset{\text{weak acid}}{H_2CO_3} \rightarrow \underset{\text{basic salt}}{Na_2CO_3} + H_2O$$

Conjugate Acid and Conjugate Base :-

Consider the reaction between HCl and H_2O

$$\underset{\text{(acid)}}{HCl} + \underset{\text{(base)}}{H_2O} \rightarrow \underset{\text{(conjugate acid)}}{H_3O^+} + \underset{\text{(conjugate base)}}{Cl^-}$$

Therefore:

a: A positively charged specie formed as a result of acceptance of proton by a base is conjugate acid (H_3O^+)

b: A negatively charged specie formed as a result of donation of proton by a an acid is called conjugate base (Cl^-).

→ Conjugate acid – base pairs:

- When an acid reacts with base, a conjugate acid and conjugate base is formed.
- The pair of acid and conjugate base **OR** the pair of base and conjugate acid is called as **conjugate acid - base pair.**

i. The species of this pair differ by H^+ ion.

ii. Weak acids have strong conjugate base-acid pair.

iii. Strong acids have weak conjugate base-acid pair.

$$A_1 + B_2 \rightarrow B_1 + A_2$$

acid base conjugate base conjugate acid

here A_1B_1 and A_2B_2 are conjugate acid-base pairs.

Examples:

1. $HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$
2. $HCl + NH_3 \rightleftharpoons Cl^- + NH_4^{+1}$
3. $HNO_3 + H_2O \rightleftharpoons NO_3^{-1} + H_3O^{+1}$
4. $NH_4^{+1} + H_2O \rightleftharpoons NH_3 + H_3O^{+1}$
5. $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$
6. $H_2O + CO_3^{-2} \rightleftharpoons OH^{-1} + HCO_3^{-1}$
7. $HClO_4 + H_2O \rightleftharpoons ClO_4^{-1} + H_3O^{+1}$

All the pairs are conjugate acid - base pairs.

→ Ionization Constant of water :-

Water is amphoteric substance which can act as both acid or base.

(1) $$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

This reaction is called as auto ionization of water.

- One molecule of H_2O acts as an acid and donates H^+ ions. This molecule then become OH^-
- Other molecule of H_2O accepts that H^+ and becomes a base. This, then converts to H_3O^+.
- H_3O^+ is called as conjugate acid. OH^- is called as conjugate base.
- The two conjugate acid - base pairs are as follows,

$$\boxed{H_2O - OH^-}$$
$$\boxed{H_2O - H_3O^+}$$

When there is an equilibrum between H_2O, H_3O^+ and OH^-, K_c is written as :

$$K_c = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

Since water is a solvent, so its concentration remains constant. i.e.,

$$K_c [H_2O]^2 = [H_3O^+][OH^-]$$

as $K_c [H_2O]^2$ is product of two constants,
where $K_c [H_2O]^2 = K_w$
or,

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

where K_w = ion product of water

② $$H_2O \rightleftharpoons H^+ + OH^-$$

This is the simple ionization of water
At equilibrium,

$$K_c = \frac{[H^+][OH^-]}{[H_2O]}$$

As water is solvent,

$$K_c [H_2O] = [H^+][OH^-]$$

or $$K_w = [H^+][OH^-]$$

③ i. At 25°C, the K_w for pure water is

$$K_w = [H^+][OH^-] = 1\times10^{-14}$$

and as $[H^+] = [OH^-]$ so,

$$[H^+] = [OH^-] = 1\times10^{-7} \text{ each}$$

ii. At 40°C, K_w for pure water is

$$K_w = [H^+][OH^-] = 3.8\times10^{-14}$$

where $[H^+] = [OH^-] = 1.9\times10^{-7}$ each

So, K_w is **temperature dependent**.

Example: 8.1

The concentration of $[OH^-]$ in ammonia solution is $0.005M$. Calculate concentrate of $[H^+]$ in it.

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

so $[OH^-][H^+] = 1 \times 10^{-14}$

$$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]}$$

$$[H^+] = \frac{1 \times 10^{-14}}{0.005} M$$

$$[H^+] = 2 \times 10^{-12} M$$

Example: 8.2

Calculate the pH of 0.001 M HCl solution.

$$HCl \xrightleftharpoons{H_2O} H^+ + Cl^-$$

$[HCl] = [H^+] = [Cl^-]$

so $[H^+] = 0.001 M$

here $pH = -\log[H^+]$

$pH = -\log(0.001)$

$pH = 3.00$

Example: 8.3

Calculate the pH of 0.062 M NaOH solution.

$$NaOH \xrightleftharpoons{H_2O} Na^+ + OH^-$$

$[NaOH] = [Na^+] = [OH^-]$

so $[OH^-] = 0.062M$

here $pOH = -\log [OH^-]$

$pOH = -\log (0.062)$

$pOH = 1.2076$

but $pH + pOH = 14$

$pH = 14 - pOH$

$pH = 14 - 1.2076$

$pH = 12.79$

Activity: 1 (Self Check: 8.1)

What is the pH of solution containing 1.95g pure H_2SO_4 per dm^3 of solution?

i) Concentration of H_2SO_4

as $\text{concentration} = \frac{\text{moles}}{\text{volume}}$

$\text{moles} = \frac{\text{mass in g}}{\text{molar mass}} = \frac{1.95g}{98gmol^-} = 0.019$ moles

so $\text{conc.} = \frac{0.019 \text{ mol}}{1 dm^3} = 0.019 M$

$[H_2SO_4] = 0.019M$

ii) pH calculation

as $H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{-2}$

so conc of H^+ = 2(conc. of H_2SO_4)

i.e $[H^+] = 2 \times 0.019 M$

$[H^+] = 0.0396 M$

but $pH = -\log [H^+]$
$pH = -\log (0.0396)$
$pH = 1.4$

note: The whole topic of titration is discussed in experiments.

→ Strong and Weak acid.

According to Bronsted, acid is a substance that donates proton. The acid which has high tendency of donating proton is a strong acid and vice versa.

Ways to recognise strength of acid:-

a) pH

pH is defined as negative log of concentration of H^+.

$$pH = \frac{1}{\log [H^+]} = -\log [H^+]$$

if $pH = 7$, neutral solution
if $pH < 7$, acidic solution
if $pH > 7$, basic solution

Lesser is the pH, stronger is the acid.

b) % ionization

It is given as,

$$\%\ \text{ionization} = \frac{\text{amount of acid ionized}}{\text{total amount of acid}} \times 100$$

The acid with highest % ionization would be strongest and vice versa.

c) Ka

Ka is the ionization constant of acid.
Suppose reaction with general acid "**HX**".

$$HX + H_2O \rightleftharpoons H_3O^+ + X^-$$

Kc at equilibrium is

$$K_c = \frac{[H_3O^+][X^-]}{[HX][H_2O]}$$

as ∴ water is solvent and has constant concentration so,

$$\therefore K_c\,[H_2O] = \frac{[H_3O^+][X^-]}{[HX]}$$

The product of two constants Kc and $[H_2O]$ is equal to ionization constant of acid, Ka.

so

$$\mathbf{K_a = \frac{[H_3O^+][X^-]}{[XH]}}$$

Ka is a temperature dependent quantity.

"Higher is value of Ka, greater is the strength of acid."

if $K_a > 1$ strong acid
if $K_a = 1 \rightarrow 10^{-3}$ moderate acid
if $K_a < 10^{-3}$ weak acid

d) pKa

pKa is defined as (-) negative log of Ka.

$$pK_a = -\log K_a$$

Smaller is the value of pKa, large will be Ka and stronger would be the acid.

→ **Strong and weak base**

According to Bronsted, base is a proton acceptor. The base which has high tendency to accept proton is stronger and vice versa.

Ways to recognise strength of base:-

a) pOH

pOH is defined as negative log of concentration of OH^- ions.

$$pOH = \frac{1}{\log[OH^-]} = -\log[\bar{O}H]$$

if pOH = 7 neutral solution
if pOH > 7 basic "
" " < 7 acidic "

Stronger bases have high value of pOH.

b) Kb

Kb is defined as ionization constant of base.

Suppose a reaction with general base **B**

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

Kc at equilibrium is

$$K_c = \frac{[BH^+][OH^-]}{[B][H_2O]}$$

As water is a solvent, so its concentration is constant

$$K_c[H_2O] = \frac{[BH^+][OH^-]}{[B]}$$

here $K_c[H_2O]$ is product of two constants equal to ionization constant of base K_b.

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

Greater is the value of K_b, stronger is the base.

c) pK_b

pK_b is defined as negative log of K_b.

$$pK_b = -\log K_b$$

Smaller is the value of pK_b, larger will be the K_b and stronger is the base.

Short Proofs

i) $pH + pOH = 14$

ionic product of water at 25°C

$$K_w = [H^+][OH^-] = 1\times10^{-14}$$

→ applying log

$$\log[H^+][OH^-] = \log 1\times10^{-14}$$

$$\log[H^+] + \log[OH^-] = -14\log 10$$

→ multiplying by (−1)

$$(-\log[H^+]) + (-\log[OH^-]) = +14\log 10$$

$$pH + pOH = 14 \qquad \because \log 10 = 1$$

ii) $K_a \propto \frac{1}{K_b}$

▸ $HA \xrightleftharpoons{H_2O} H^+ + A^-$

$$K_a = \frac{[H^+][A^-]}{[HA]} \quad \text{— i}$$

▸ $A^- + H_2O \rightleftharpoons HA + OH^-$

$$K_c = \frac{[HA][OH^-]}{[A^-][H_2O]}$$

$$K_c[H_2O] = \frac{[HA][OH^-]}{[A^-]}$$

$$K_b = \frac{[HA][OH^-]}{[A^-]} \quad \text{— ii}$$

▸ multiplying i and ii

$$K_a \times K_b = \frac{[H^+][\cancel{A^-}]}{[H\cancel{A}]} \times \frac{[H\cancel{A}][OH^-]}{[\cancel{A^-}]}$$

$$K_a \times K_b = [H^+][OH^-]$$

$$K_a \times K_b = K_w \quad \text{— iii}$$

$$K_a = \frac{K_w}{K_b}$$

taking Kw as constant.

$$K_a \propto \frac{1}{K_b}$$

iii) $pK_a + pK_b = pK_w$

as mentioned above in iii

$$K_a \times K_b = K_w$$

▸ applying log

$$\log(K_a \cdot K_b) = \log K_w$$

$$\log K_a + \log K_b = \log K_w$$

- multiplying both sides by (-1)

$$-\log Ka - \log Kb = -\log Kw$$

or $\underline{pKa + pKb = pKw}$

(iv) $pKa + pKb = 14$

as $Ka \times Kb = Kw$

$\therefore Kw = 1 \times 10^{-14}$

$Ka \times Kb = 1 \times 10^{-14}$

- applying log on both sides

$$\log Ka \times Kb = \log 10^{-14}$$

$$\log Ka + \log Kb = -14 \log 10 \quad \because \log 10 = 1$$

- multiplying (-1) with both sides

$$-\log Ka - \log Kb = (-14)(-1)$$

$$\underline{pKa + pKb = 14}$$

Example 8.6

Calculate concentration of H^+ of solution having 1.0 M HF $(Ka = 7.2 \times 10^{-4})$

$$HF \rightleftharpoons H^+ + F^-$$

so $[HF] = [H^+] = [F^-]$

	$HF \rightleftharpoons$	H^+	$+ F^-$
$t = 0$	1.0	0	0
$t = t_{eq}$	$1.0 - x$	x	x

$$Ka = \frac{[H^+][F^-]}{[HF]}$$

$$7.2 \times 10^{-4} = \frac{x^2}{1.0 - x}$$

taking square root on both sides

$$\sqrt{7.2\times10^{-4}} = \frac{\sqrt{x^2}}{\sqrt{1-x}}$$

$$0.0268 = \frac{x}{\sqrt{1-x}}$$

as x is very small as compared t
1, so $x \approx 0$.

$$\therefore 0.0268 = \frac{x}{\sqrt{1-0}}$$

$$0.0268 = x$$

$$[H^+] = x = 0.0268\ M$$

Example : (Class Activity)

The pH of 0.1M solution of an acid is 2.85. Find Ka for acid.

Suppose :
Acids donate a proton (H^+) and a conjugated base (A^-)

$$HA \rightleftharpoons H^+ + A^-$$

here $pH = 2.85$ $\quad\quad \therefore pH = -\log[H^+]$

$[H^+] = -\log^{-1} pH$ $\quad\quad \therefore [H^+] = -\log^{-1} pH$

$$[H^+] = -\text{antilog } 2.85$$

$$[H^+] = 1.41\times10^{-3}$$

initial conc. of $[HA] = 0.1M$

	acid $\rightleftharpoons$	H^+ +	conjugate base
$t=0$	0.1	0	0
$t=eq$	$0.1-1.41\times10^{-3}$ $= 0.098$	1.41×10^{-3}	1.41×10^{-3}

$$Ka = \frac{[H^+][A^-]}{[HA]}$$

$$Ka = \frac{[1.41\times10^{-3}][1.41\times10^{-3}]}{[0.098]}$$

$$Ka = \frac{(1.41\times10^{-3})^2}{0.098}$$

$$Ka = 2.01\times10^{-5}$$

Example : 8.7

pKa of acetic acid at 25°C is 4.76. Calculate pKb of conjugate base of acetic acid.

Using relation,

$$pKa + pKb = 14$$
$$pKb = 14 - pKa$$
$$pKb = 14 - 4.76$$
$$pKb = 9.24$$

Example : 8.8

pKb of pyridine at 25°C is 8.25. Calculate pKa of conjugate acid of pyridine.

Using relation,

$$pKa + pKb = 14$$
$$pKa = 14 - pKb$$
$$pKa = 14 - 8.25$$
$$pKa = 5.75$$

Activity (Class activity)

Example :

What is the percentage ionisation of acetic acid in solution in which 0.1M of it has been dissolved per dm^3 of solution?

$(Ka = 1.8 \times 10^{-5})$

CH_3COOH ionizes to produce CH_3COO^{-1} and H^{+1} ions.

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$

$t = 0$	$0.1M$	0	0
$t = t_{eq}$	$0.1 - x$	x	x

$$Ka = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$Ka = \frac{x^2}{0.1 - x} \quad (\text{where } x \approx 0)$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.1}$$

$$\sqrt{1.8 \times 10^{-5}} \cdot \sqrt{0.1} = \sqrt{x^2}$$

$$x = 1.34 \times 10^{-3}$$

conc. of acid ionized $= 0.1 - x$

$= 0.1 - 1.34 \times 10^{-3}$

$= 0.098$

% ionization $= \frac{\text{conc. of acid ionized}}{\text{initial conc.}} \times 100$

$= \frac{0.098}{0.1} \times 100$

% ionisation $= 98\%$

Lewis Definitions of acid and base

In 1932, Lewis gave definition of acids and bases on basis of electron pair.

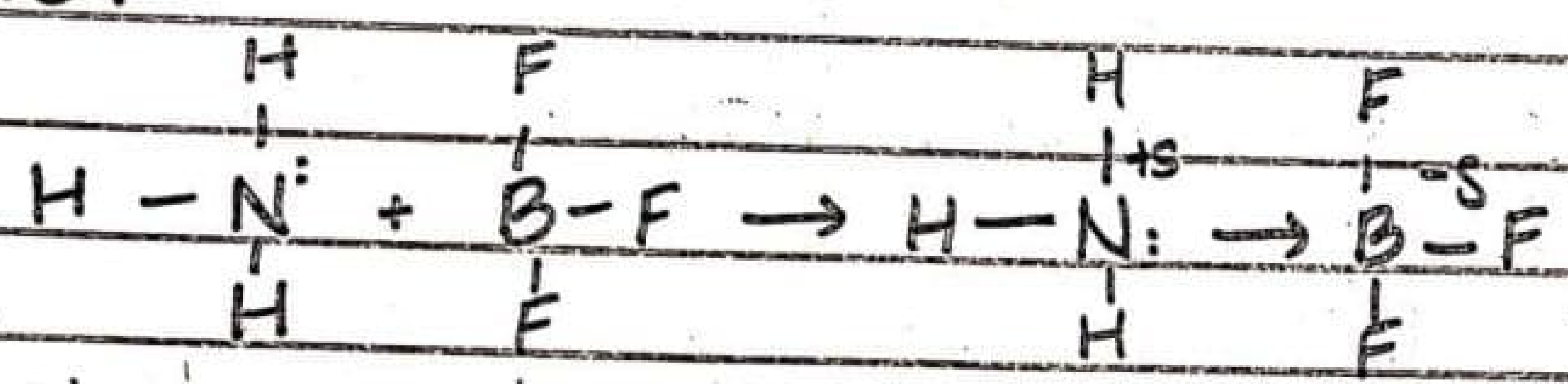
Lewis Acid:

An acid is a substance which can accept a pair of electrons. So, lewis acid has no unpaired electrons on central atom.

Lewis Base:

A base is a substance, which can donate a pair of electrons. So, lewis base has a pair of electrons on central atom.

Example:



In this reaction, ammonia donates a pair of electron to BF_3. So, NH_3 is base and BF_3 is acid.

The arrow → shows a coordinate covalent bond between lewis acid and lewis base.

Importance of Lewis concept :-

Lewis concept of acid and base is more general and easily explains those acids and bases which are not explained by Bronsted base acid concept.

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"Buffer Solution"

i. Definition:-

Solution whose pH do not change by addition of small amount of acid, base, or dilution or keeping it for a long time, is called buffer solution.

ii. Types :-

There are two types of buffer solution

→ **Acidic Buffer Solution**

It is the mixture of weak acid and its salt with strong base is called acidic buffer solution. Its pH is less than 7.

- $CH_3COOH + CH_3COONa$
- $H_3PO_4 + Na_3PO_4$

→ **Basic Buffer Solution**

It is the mixture of weak base and its salt with strong acid is called basic buffer solution. Its pH is greater than 7.

- $NH_4OH + NH_4Cl$
- $Al(OH)_3 + AlCl_3$

ii. **Buffer Action:-**

Buffer solution has a constant value of pH. The method by which a buffer can maintain its pH is called buffer action

(1)

Acidic Buffer solution is ionized as,

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$
$$CH_3COO^-Na \rightleftharpoons CH_3COO^- + Na^+$$

i) On adding acid

When acid is added in buffer solution, H^+ of acetic acid combine with acetate ion (CH_3COO^-) and form acetic acid which is already present in solution so there is no change in pH.

$$CH_3COO^- + H^+ \rightarrow CH_3COOH$$

ii) On adding base

When acidic buffer solution gets some base, the OH^- of base combine with H^+ of buffer to form water which is neutral, so there is no change in pH

$$H^+ + OH^- \rightarrow H_2O$$

(2)

Basic Buffer solution is ionized as,

$$NH_4OH \rightleftharpoons NH_4^+ + OH^-$$
$$NH_4Cl \rightleftharpoons NH_4^+ + Cl^-$$

i) On adding acid

When acid is added in basic buffer solution, H^+ of acid combine with OH^- ion and form water which is neutral. So pH do not change.

$$H^+ + OH^- \rightarrow H_2O$$

ii) On adding base

When base is added in buffer solution, OH^- of base combine with NH_4^+ to form NH_4OH which is already present in solution, so there is no pH change.

$$NH_4^+ + OH^- \rightarrow NH_4OH$$

iv. **Facts about buffer solution :-**

1) When a weak acid and its conjugate base is added in water OR when a weak base and its conjugate acid is added in water, a buffer solution is formed.
2) Buffer solution always resists change in pH and hence its pH is constant.
3) Whenever the acid or base is added in buffer, the H^+ or OH^- ions are produced which react with negative and positive ions of buffer and hence pH is made constant.
4) When a strong base or strong acid

is used, the buffer solution is not formed.

5) Examples of buffer are:

- $CH_3COOH + CH_3COONa$
- $H_3PO_4 + Na_3PO_4$
- fomic acid + sodium format

V. Applications of buffer :-

i Industrial Application:

They are used in industries to prepare photographic material, leather and dyes.

ii Use in Processes:

They are used in process of electro-plating and analytical procedures like titration.

iii pH measurement:

The buffer solutions are used as a standard for calibration of pH meters.

iv Preparation of Culture:

Buffer is added in culture of bacteria so that pH of culture remains constant.

vi. Natural Buffer Solutions :-

human blood (7.35 — 7.45)
tears (7.40)
stomach juice (1.65 – 1.75)
milk (6.7 – 6.8)
egg white (8.0 – 8.1)

Example 8.9 :

a) Calculate pH of acetic acid + sodium acetate buffer solution containing 1mol each.

CH_3COOH ionizes as,

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$

$$Ka = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$1.8 \times 10^{-5} = Ka$$

$pH = ?$

$CH_3COOH = 1M$

$CH_3COONa = 1M$

$$pKa = -\log 1.8 \times 10^{-5}$$
$$= 4.74$$

$$pH = pKa + \log \frac{[CH_3COONa]}{[CH_3COOH]}$$

$$pH = pKa + \log \frac{1}{1}$$

$$pH = pKa + \log 1$$

$$\log 1 = 0$$

$$pH = pKa + 0$$

$$pH = 4.74$$

b) When 0.01 mole of HCl is added,

pH = ?

$K_a = 1.85 \times 10^{-5}$

$pK_a = -\log 1.8 \times 10^{-5} = 4.74$

When 0.01 mol of HCl is added,

new concentration;

$[CH_3COOH] = 1\,mol + 0.01\,mol$

$= 1.01\ mol\ dm^{-3}$

$[CH_3COONa] = 1\,mol - 0.01\,mol$

$= 0.99\ mol\ dm^{-3}$

$$pH = pK_a + \log \frac{[CH_3COONa]}{[CH_3COOH]}$$

$$pH = 4.74 + \log \frac{0.99}{1.01}$$

$$pH = 4.736$$

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Example: 10

What is the pH of buffer if concentration of CH_3COOH is $0.1\ dm^{-3}$ and CH_3COONa is $1.0\ M\ dm^{-3}$.

$pKa = 4.76$ for CH_3COOH

$$pH = pKa + \log \frac{[salt]}{[acid]}$$

$$= 4.76 + \log \frac{(1.0)}{(0.1)}$$

$$= 4.76 + \log 10$$

$$= 4.76 + 1$$

$$pH = 5.76$$

Activity (Self Check, 8.4)

Calculate pH of buffer in which 0.11 M CH_3COONa and 0.09 M CH_3COOH is present $Ka = 1.8 \times 10^{-5}$

$$pKa = -\log Ka$$

$$= -\log(1.8 \times 10^{-5})$$

$$= 4.745$$

$$pH = pKa + \log \frac{[salt]}{[acid]}$$

$$= 4.76 + \log \frac{(0.11)}{(0.09)}$$

$$= 4.76 + \log 1.22$$

$$= 4.76 + 0.087$$

$$pH = 4.84$$

Henderson's Equation for pH

Prove $pH = pKa + \log \frac{[salt]}{[acid]}$

$$HA \rightleftharpoons H^+ + A^-$$

Ka for HA

$$Ka = \frac{[H^+][A^-]}{[HA]}$$

$$[H^+] = \frac{Ka[HA]}{[A^-]}$$

where $[HA] = [acid]$

$[A^-] = [salt]$

then,

$$[H^+] = Ka \frac{[Acid]}{[salt]}$$

$$\log [H^+] = \log\left(Ka \cdot \frac{[Acid]}{[salt]}\right) \quad \because \log(a \cdot b) = \log a + \log b$$

$$\log [H^+] = \log Ka + \log \frac{[Acid]}{[salt]}$$

$$\log [H^+] = \log Ka + \log [Acid] - \log [salt]$$

or,

$$-\log [H^+] = -\log Ka - \log [Acid] + \log [salt]$$

$$pH = pKa + \log [salt] - \log [acid]$$

$$pH = pKa + \log \frac{[salt]}{[acid]} \quad \because \log a - \log b = \log (a/b)$$

for pOH,

$$pOH = pKb + \log \frac{[salt]}{[base]}$$

Salt Hydrolysis

i Hydrolysis is the break down of water into H^+ and OH^- ions.

ii Salt hydrolysis is defined as a reaction between cations and anions of salts with H^+ and OH^- of water.

Generalized Salt Hydrolysis

Consider a salt **MX** dissolved in water. It produces M^+ and X^- ions.

"$MX \rightarrow M^+ + X^-$"

The M^+ reacts with OH^- of water as follows.

"$M^+ + H_2O \rightarrow MOH + H^+$"

The X^- reacts with H^+ of water as follows.

"$X^- + H_2O \rightarrow HX + OH^-$"

Four types of Salt Hydrolysis

P.T.O

(i)

When the salts of weaks acids or strong bases hydrolyse, basic solution is produced. pH of solution is greater that 7.

e.g. $CH_3COONa + H_2O \rightarrow CH_3COOH + NaOH$

$\underset{\text{weak acid}}{\text{salt of}} + H_2O \rightarrow \underset{\text{acid}}{\text{weak}} + \underset{\text{base}}{\text{strong}}$

So overall solution is basic.

(ii)

When the salts of strong acids or weak bases hydrolyse in water, acidic solution is produced with pH < 7.

e.g. $CuSO_4 + 2H_2O \rightarrow Cu(OH)_2 + H_2SO_4$

$\underset{\text{strong acid}}{\text{salt of}} + 2H_2O \rightarrow \underset{\text{base}}{\text{weak}} + \underset{\text{acid}}{\text{strong}}$

So overall, solution is acidic.

(iii)

When salt of strong acid and strong base are hydrolysed, they give stron acids and bases. So the solution becomes neutral with pH = 7.

e.g. $NaCl + H_2O \rightarrow NaOH + HCl$

$$\underset{\text{salt of strong acid}}{} + H_2O \rightarrow \underset{\text{base}}{\text{strong}} + \underset{\text{acid}}{\text{strong}}$$

So the strong base and strong acid neutralize each other and pH of solution becomes 7.

(iv)

Salts of weak acids and weak bases hydrolyse but the resulting solution can be neutral, acidic or basic as follows.

if, $K_a = K_b$ neutral

$K_a > K_b$ acidic

$K_a < K_b$ basic

Salt Type: Acid	Base	Ion which imparts hydrolysis	Solution pH
strong	strong	none	7
strong	weak	cations	<7
weak	strong	anions	>7
weak	weak	cation + anions	may be =/</> 7

levelling Effect

Definition:-

If we use a base to check strength of different acids but the base shows tha all acids are of same strength, then thi inability of base is called levelling effect as the base shows the strength of all acids at equal level.

Example:-

Consider water as a base, then its shows acidic strengths as follows.

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

$$HI + H_2O \rightarrow H_3O^+ + I^-$$

$$HBr + H_2O \rightarrow H_3O^+ + Br^-$$

$$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$$

As all the acids have donated one H^+, the conjugate acids thus formed are al H_3O^+ which shows that the strength of all acids is same. This is levellir effect of water.

We can use CH_3COOH as base inste

of H_2O to check strengths of acids.

date

For video Lectures visit you tube

Chemistry by Prof. Javed Iqbal

The End

Ch#8 "EXERCISE"

Q1: MCQ's

(i)	a	(ii)	c	(iii)	b
(iv)	d	(v)	b	(vi)	d
(vii)	b	(viii)	d	(ix)	b
(x)	d	(xi)	c	(xii)	b
(xiii)	c	(xiv)	a	(xv)	b

Q2: Short Questions:-

★ **Acidic Substance :-**

"A substance which give acidic solution when dissolved in water."

i.e: Acids turn blue litmus red.

Oxides of non-metals: e.g:- CO_2, SO_2 & N_2O_5 produce H_2CO_3, $2HNO_3$ when dissolved in water.

e.g: $CO_2 + H_2O \longrightarrow H_2CO_3$

★ **Basic Substance :-**

"A substance which gives basic solution when dissolve in water.

Oxides of metals:- like CaO & Na_2O when dissolved in water produce $NaOH$ and $Ca(OH)_2$

e.g: $Na_2O + H_2O \rightleftharpoons 2NaOH$

Amphoteric substance:

"A Substance when dissolve in water produce such solution which are of amphoteric nature." i.e:- "Such substance behave an acid as well as base"

Less electropositive elements are amphoteric substance. e.g: Al_2O_3

$$Al_2O_3 + 6HCl \rightleftharpoons 2AlCl_3 + 3H_2O \text{ (as base)}$$

$$Al_2O_3 + 6HCl \rightleftharpoons 2NaAlO_2 + H_2O \text{ (as acid)}$$

Q. Extra — **Amphoteric substance/substance having amphoteric nature:**

"There are substances like H_2O and CH_3COOH which act as both acid and base, such substances are called as amphoteric" i.e: Some substances when dissolved in water produce such solution which are of amphoteric nature.

* **Less electropositive** elements are amphoteric in nature.

e.g: H_2O:-

$$H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO_4^- \text{ (as base)}$$

$$NH_3 + H_2O \longrightarrow NH_4^+ + OH^- \text{ (as acid)}$$

CH_3COOH:-

$$CH_3COOH + H_2O \longrightarrow CH_3COO^- + H_3O^+ \text{ (as acid)}$$

$$CH_3COOH + HCl \longrightarrow CH_3COOH_2^+ + Cl^- \text{ (as base)}$$

(ii) Bronsted-Lowery Acid theory:-

Bronsted Acid is anything that donates proton or hydrogen ion (H^+)

Example:

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

In this reaction the HCl donate proton to water and becomes Cl^-. So, it is an acid here while water would be base.

Bronsted-Lowery base theory:-

Bronsted base is anything that accept proton or hydrogen ion (H^+).

Example:-

i. $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$

In this reaction water accepts proton or (H^+) and becomes H_3O^+. So it acts as a base.

(iii) Conjugate acid-base pairs:-

* When an acid react with base, conjugate acid and conjugate base is formed.
* The pair of acid and conjugate base.
  The pair of base and conjugate acid is known as Conjugate-acid-base pairs.

i: The specie of this pair is differ by ion.
ii: Weak acids have strong conjugate base-acid pair.
iii: Strong acids have weak conjugate base-acid pair.

$$A_1 + B_2 \longrightarrow B_1 + A_2$$

acid base conjugate base conjugate acid

here A_1B_2 and B_1A_2 are conjugate acid-base pair.

Examples:

$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$
$$HCl + NH_3 \rightleftharpoons Cl^- + NH_4^+$$
$$HNO_3 + H_2O \rightleftharpoons NO_3^- + H_3O^+$$
$$NH_4 + H_2O \rightleftharpoons H_3O^{+1} + NH_3$$
$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$
$$H_2O + CO_3^{-2} \rightleftharpoons OH^- + HCO_3^{-1}$$
$$HClO_4 + H_2O \rightleftharpoons ClO_4^- + H_3O^{+1}$$

All the pairs are conjugate acid-base pairs.

(iv) Lewis difinitions of acid and base:-

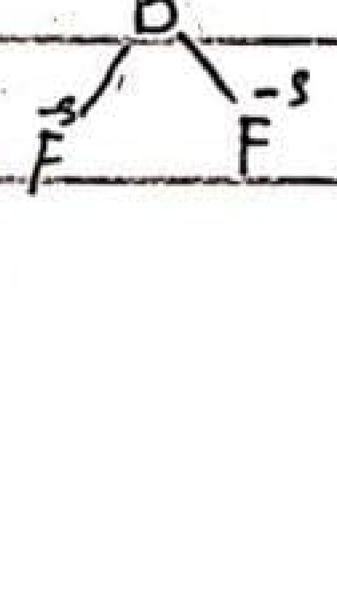
In 1932, Lewis gave difinition of acids and bases on basis of electron pair.

* Lewis Acid:

An acid is a substance which can accept a pair of electrons. So, Lewis acid has no unpaired electron on central atom.

* It is also called **electrophile** (electron defficient

Example: BF_3

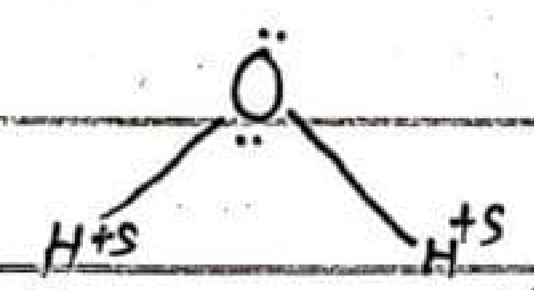


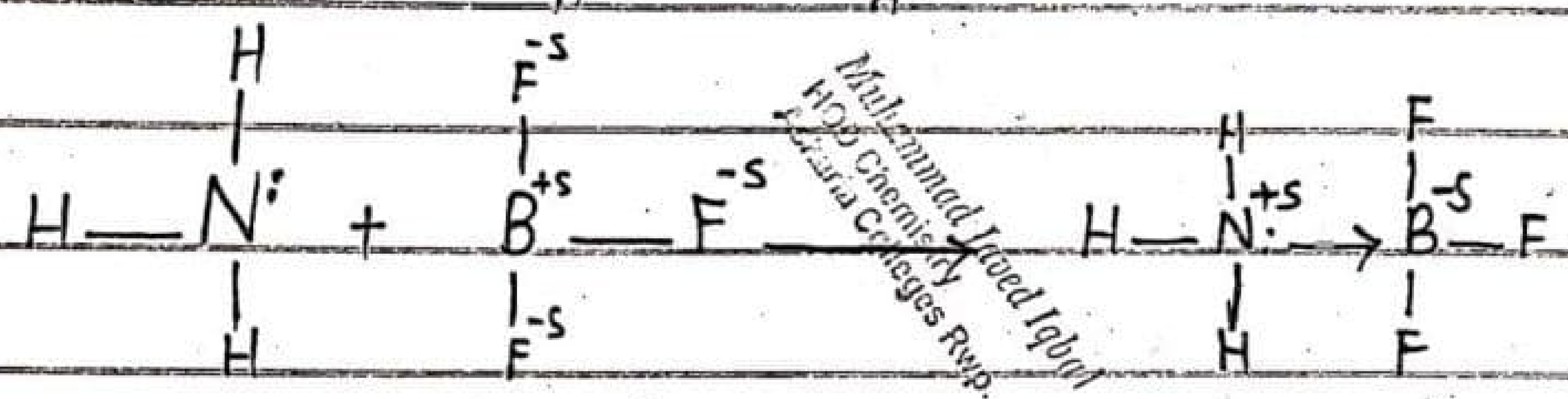
★ Lewis Base:-

A base is a substance which can donate a pair of electrons. So, Lewis base has a pair of electrons on central atom.

★ It is also called nucleophile (electron rich)

Example: H_2O





Base **Acid**

In this reaction ammonia donates a pair of electron to BF_3. So, NH_3 is base and BF_3 is acid.

The arrow ($\longrightarrow$) shows a coordinate covalent bond b/w lewis acid and lewis base

(V.) Classification of bronsted acid or base:-

(i) HCO_3^{-1} : Base

(ii) HBr : Acid

(iii) CH_3COO^{-1} :- Base

(vi) **Gastric acidity:→**
Hydrocloric acid (HCl) is present in our stomach, and work in specific limit to digest food. When we eat spicy food acidity of HCl becomes over in stomach.

Anti-Acid drugs:-
We use weak bases like $Mg(OH)_2$ and $Al(OH)_3$ to decrease acidity of HCl in stomach is known as anti-acid drugs.

(vii) **Ionization of water:-**
Water is amphoteric substance which can act as both acid or base.

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

1. This reaction is called as auto ionization of water.
2. One molecule of H_2O acts as an acid and donates H^+ ions. This molecule then become OH^-.
3. Other molecule of H_2O acepts that H^+ and becomes a base. This, then converts to H_3O^+.

VIII) pH:-

pH is defined as negative log of Concentration of H^+.

$$pH = \frac{1}{\log[H^+]} \Rightarrow -\log[H^+]$$

* **Value of pH for acidic solution:-**

if,

$pH < 7$ then solution is known as acidic solution.

* **Value of pH for basic solution**

if,

$pH > 7$ then Solution is known as basic solution.

* **Value of pH for Neutral Solution:**

if, $pH = 7$

then solution is known as neutral solution.

(a) Ka :-

Ka is the ionization constant of acid

Suppose reaction with general acid "HX".

$$HX + H_2O \rightleftharpoons H_3O^+ + X^-$$

Kc at equilibrium is,

$$K_c = \frac{[H_3O^+][X^-]}{[HX][H_2O]}$$

as water is solvent and has constant concentration so,

$$K_c[H_2O] = \frac{[H_3O^+][X^-]}{[HX]}$$

The product of two constant K_c and $[H_2O]$ is equal to ionization constant of acid K_a

$$K_a = \frac{[H_3O^+][X^-]}{[HX]}$$

K_a is temperature dependent quantity

"Higher is value of K_a, greater is the strength of acid".

if $K_a > 1$ strong acid

if $K_a = 1 - 10^{-3}$ moderate acid

if $K_a < 10^{-3}$ weak acid

(b) **pKa:**

It is defined as negative (–) log of K_a.

$$pK_a = -\log K_a$$

Smaller is the value of pK_a, larger will be K_a and stronger would be the acid.

Q
Extra

Curdling of milk with Lemon juice:-

Milk is composed of many components like protein, carbohydrates, and water One of its component is known as casein. Its molecule repell eachother due to having same -ve charge So, lemon juice is added to neutralize these molecule. As a result curd is formed. this is known as curdling of milk with lemon juice.

(X) (a) Kb:-

Kb is defined as "ionization constant of base."

Suppose a reaction with general base "B".

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

Kc at equilibrium is

$$K_c = \frac{[BH^+][OH^-]}{[B][H_2O]}$$

As water is a solvent, so its concentration is constant.

$$K_c[H_2O] = \frac{[BH^+][OH^-]}{[B]}$$

here $K_c[H_2O]$ is product of two constants equal to ionization constant of base K_b.

10

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

Greater is the value of kb, strong is the base.

(b) **pkb :-**

pkb is defined as negative log of kb.

$$pkb = -\log kb$$

Smaller the value of pkb, larger will be kb and stronger is the base.

Q: **Ionized salt and its use in practical life:-**
Extra

Ionized Salt:

NaCl is ionized salt and mixture of KI and NaI.

Use:

NaCl is use in our daily life mostly as a table-salt.

Cause of deficiency:-

Deficiency of KI and NaI (iodine salt) cause many disease In which thyroid gland disease is most common, known as goiter.

Goiter:-

It is a disease in which patient's throat swells to a big size.

Relationship of Ka and Kb:-

* $Ka \propto \frac{1}{Kb}$:-

$$HA \underset{}{\overset{H_2O}{\rightleftharpoons}} H^+ + A^-$$

$$Ka = \frac{[H^+][A^-]}{[HA]} \longrightarrow (i)$$

$$A^- + H_2O \rightleftharpoons HA + OH^-$$

$$Kc = \frac{[HA][OH^-]}{[A^-][H_2O]}$$

$$Kc[H_2O] = \frac{[HA][OH^-]}{[A^-]}$$

$$Kb = \frac{[HA][OH^-]}{[A^-]} \longrightarrow (ii)$$

multiplying (i) and (ii)

$$Ka \times Kb = \frac{[H^+]\cancel{[A^-]}}{\cancel{[HA]}} \times \frac{\cancel{[HA]}[OH^-]}{\cancel{[A^-]}}$$

$$Ka \times Kb = [H^+][OH^-]$$

$$Ka \times Kb = Kw \longrightarrow (iii) \quad \because Kw = [H^+][OH^-]$$

$$Ka = \frac{Kw}{Kb}$$

taking Kw as constant.

$$\boxed{Ka \propto \frac{1}{Kb}}$$

Q.14 Concentration of slightly soluble salt:

Extra: Slightly Soluble Salt is dissolved in water like $PbSO_4$

$$PbSO_4 \rightleftharpoons Pb^{+2}_{(aq)} + SO_{4(aq)}^{-2}$$

$$K_c = \frac{[Pb^{+2}][SO_4^{-2}]}{[PbSO_4]}$$ Law of mass action

The rate of change of concentration of $PbSO_4$ is so small, that it almost remains constant.

$$K_c\,[PbSO_4] = [Pb^{+2}][SO_4^{-2}]$$ or

$$K_{sp} = [Pb^{+2}][SO_4^{-2}]$$ where $K_c[PbSO_4] = K_{sp}$

Where **Ksp** is called solubility product.

(XIII) Buffer Solution:-

"Solution whose pH do not change by addition of small amount of acid, base or dilution or keeping it for a long time is called buffer solution"

Examples:-

- $CH_3COOH + CH_3COONa$
- $H_3PO_4 + Na_3PO_4$
- formic acid + Sodium format

23 Ionization equation of water and its constant and its applications:-

Water is amphoteric substance which can act as both acid or base.

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

This reaction is called auto-ionization of water.

1. One molecule of H_2O acts as acid and donates H^+ ions. This molecule than become OH^- (hydroxide ion).

2. Other molecule of H_2O acts as base and accepts H^+ ions and become base. This is than convert into H_3O^+ (hydronium ion).
3. Two conjugate acid-base pairs are as follows.

   H_2O — OH^- (acid-conjugate base)

   H_2O — H_3O^+ (base-conjugate acid)
4. OH^- is called conjugate base.
5. H_3O^+ is called conjugate acid.

When there is equilibrium between H_2O, H_3O^+ and OH^-, kc is written as:

$$kc = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

Since, water is solvent its concentration remains constant.

i.e:

$$K_c [H_2O]^2 = [H_3O^+][OH^-]$$

as $K_c [H_2O]^2$ is product of two constan
where $K_c [H_2O]^2 = K_w$

or,-- So,

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

where K_w = ion-product constant of water

Ionization constant:

This is simple ionization of water,
At equilibrium

$$K_c = \frac{[H^+][OH^-]}{[H_2O]}$$

As water is solvent;

$$K_c [H_2O] = [H^+][OH^-]$$

or $K_w = [H^+][OH^-]$

Applications of ionization constant:

At **25°c**, the K_w for pure water is

$$K_w = [H^+][OH^-] = 1 \times 10^{-7}$$

and as $[H^+] = [OH^-]$ so,

$$[H^+] = [OH^-] = 1 \times 10^{-7} \text{ each}$$

At 40°c, K_w for pure water is

$$K_w = [H^+][OH^-] = 3.8 \times 10^{-14}$$

where $[H^+] = [OH^-] = 1.9 \times 10^{-7}$ each

So, K_w is temperature dependent

If K_w has small value than H_2O is weak electrolyte.

If kw has large value than H_2O is strong electrolyte.
With increasing temperature value of kw is increase.

Importance:-

Concentration of $[H^+]$ and $[OH^-]$ are very small numbers and inconvenient to work with, a more paractical measures calledpH.

Q4 Buffer Solutions:

"Solution whose pH do not change addition of small amount of acid, base or keeping it for a long time."

Types:

There are two types of buffer solution.

→ **Acidic buffer solution:-**

It is the mixture of weak acid and its salt with strong base is called acidic buffer solution. Its pH is less than 7.

- $CH_3COOH + CH_3COONa$
- $H_3PO_4 + Na_3PO_4$
- $HCOOH + HCOONa$

→ **Basic buffer Solution:-**

It is the mixture of weak base and its salt with strong acid, is called basic buffer solution. Its pH is always great than 7.

- $NH_4OH + NH_4Cl$
- $Al(OH)_3 + AlCl_3$

→ **Buffer action:-**

Method in which buffer solution has constant value of pH. The method by which a buffer can maintain its pH called buffer action.

→ **Acidic Buffer Solution ionization:-**

Acidic Buffer Solution is ionized as:

$$CH_3COOH \longrightarrow CH_3COO^- + H^+$$

$$CH_3COONa \longrightarrow CH_3COO^- + Na^+$$

i) **On adding acid:**

When acid is added in buffer solution H^+ acetic acid combine with acetate ion(CH_3 and form acetic acid which is already pre solution so there is no change in

e.g: $CH_3COO^- + H^+ \longrightarrow CH_3COOH$

ii) **On adding base:-**

When acidic buffer solution gets som base, the OH^- of base combine with H^+ of buffer to form water which is neutral, so there is no change in pH

e.g:- $H^+ + OH^- \longrightarrow H_2O$

Facts about buffer Solutions:

1: When a weak acid and its conjugate base is added in water OR when a weak base and its conjugate acid is added in water, a buffer Solution is formed.

2: Buffer Solution always resists change in pH and hence its pH is constant.

3: Whenever the acid or base is added in buffer, the OH^- or H^+ ions are produced which react with negative and positive ions of buffer and hence pH is made constant.

4: When a strong base or strong acid is used the buffer solution is not formed.

5: Examples of buffer are:

$CH_3COOH + CH_3COONa$

$H_3PO_4 + Na_3PO_4$

formic acid + Sodium formate butter

→ **Applications of buffer solutions:**

i: **Industrial Application:**

They are used in industries to prepare photographic material, leather and dyes.

2: **Use in Process:**

They are used in process of electro-plating and analytical procedures like titration.

iii pH measurement:-
The buffer solutions are used as standard for calibrations of pH meters.

Q5 a **Conjugate acid base pairs:**
* When an acid react with base a conjugate acid and conjugate base is formed.
* The pair of acid and conjugate base OR The pair of base and conjugate acid is called as **conjugate acid-base pairs**
i:- The specie of this pair is differ by ion.
ii: Weak acids have strong conjugate base-acid p
iii: Strong acids have weak conjugate base-acid pai

$$\underset{\text{acid}}{A_1} + \underset{\text{base}}{B_2} \longrightarrow \underset{\text{Conjugate base}}{B_1} + \underset{\text{Conjugate acid}}{A_2}$$

here A_1B_2 and B_1A_2 are conjugate acid-base p

Examples:-

1 $HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$

2 $HCl + NH_3 \rightleftharpoons NH_4^+ + Cl^-$

3 $HNO_3 + H_2O \rightleftharpoons NO_3^- + H_3O^+$

4 $NH_4^{+1} + H_2O \rightleftharpoons NH_3 + H_3O^{+1}$

5 $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$

6 $H_2O + CO_3^{-2} \rightleftharpoons OH^- + HCO_3^{-1}$

7 $HClO_4 + H_2O \rightleftharpoons ClO_4^- + H_3O^+$

All the pairs are conjugate acid-base pairs.

b **pka:-**

It is defined as negative Log of ka.

$$pka = -Log\ ka$$

Smaller is the value of pka, larger will be ka and stronger would be acid

c) **pkb:-**

It is defined as negative Log of kb.

$$pkb = -Log\ kb$$

Smaller the value of pkb, larger will be kb and Stronger would be the base.

Relative strength of acid:

★ **Acidic Strength:**

Ability of acid to donate proton, is called acidic strength.

Strong acid:-

Some acid can donate proton to a higher degree than another acid is said to be relatively strong acid.

e.g: HCl is relatively stronger than acetic acid. Acetic acid (CH_3COOH) is relatively stronger than H_2O.

Q6 **Hydrolysis:-**

"Hydrolysis is break down of water H^+ and OH^-". OR

" The reaction of cations and anions of Salts with water is called hydrolysis.

Explaination:-

Hydrolysis is different from hydration. In h $H-OH$ bond is broken whereas in hydration wa molecule adds up to a substance without bo breakage.

Generalized Salt hydrolysis:-

Consider a salt MX dissolved in water. It produces M^+ and X^- ions.

$$MX \longrightarrow M^+ + X^-$$

The M^+ reacts with OH^- of water as follows.

$$M^+ + H_2O \longrightarrow MOH + H^+$$

The X^- reacts with H^+ of water as follows.

$$"X^- + H_2O \longrightarrow HX + OH^-"$$

(a) K_2CO_3

$$K_2CO_3 + 2H_2O \longrightarrow H_2CO_3 + 2KOH$$

(b) NH_4Cl:

$$NH_4Cl + H_2O \longrightarrow \underset{\text{weak base}}{NH_4OH} + \underset{\text{Strong acid}}{HCl}$$

(C) $NaNO_3$

$$NaNO_3 + H_2O \longrightarrow NaOH + HNO_3$$

Q7 (a)

Solution:-

$HCOOH = 1M$

$[HCOONa] = 1M$

$pH = ?$

$Ka = 1.8 \times 10^{-4}$

$pka = -\log Ka$

$pka = -\log 1.8 \times 10^{-4}$

$pka = 3.74$

$pH = pka + \log \frac{[HCOONa]}{[HCOOH]}$

$pH = pka \log (\frac{1}{1})$

$\boxed{pH = 3.74}$

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(b)

$[HCOOH] = 1M$

$[HCOONa] = 1M$

$Ka = 1.8 \times 10^{-4}$

$pKa = 3.74$

When 0.10M HCl is added

$pH = ?$

New concentration

$[HCOOH] = 1.0 + 0.10 = 1.1M$

$[HCOONa] = 1.0 - 0.10 = 0.9M$

$pH = pka + \log \frac{[HCOONa]}{[HCOOH]}$

$pH = 3.74 + \log\left(\frac{0.9}{1.1}\right)$

$pH = 3.74 - 0.087$

$\boxed{pH = 3.65}$

Q8 (a)

Given data:

$pH = 10.6$, $H^+ ion = ?$

Solution:-

$pH = -\log H^+$

$H^+ = \text{antilog}(-pH)$

$H^+ = \text{antilog}(-10.6)$

$H^+ = 2.5 \times 10^{-11}$ moles/dm^3

(b)

Given data:-

$[H_3\overset{+}{O}] = 1.0 \times 10^{-9}$ moles/dm^3, $pOH = ?$

Solution:- $[\overset{+}{H}]$ or $[H_3\overset{+}{O}]$

$pH = -\log[H^+]$

$pH = -\log(1.0 \times 10^{-9})$

$pH = 9$

$pH + pOH = 14$

$pOH = 14 - pH$

$pOH = 14 - 9 = 5$

$\boxed{pOH = 5}$

9 (a)

Ka (dissociation constant):-

Ka is the ionization constant of acid. Suppose reaction with general acid "HX".

$$HX + H_2O \rightleftharpoons H_3O^+ + X^-$$

Kc at equilibrium is

$$Kc = \frac{[H_3O^+][X^-]}{[HX][H_2O]}$$

as water is solvent and has constant concentration

So,

$$Kc[H_2O] = \frac{[H_3O^+][X^-]}{[HX]}$$

The product of two constant Kc and $[H_2O]$ is equal to ionization constant of acid Ka.

$$Ka = \frac{[H_3O^+][X^-]}{[HX]}$$

Ka is temperature dependent quantity.

"Higher is value of Ka, greater is the Strength of acid."

if ka >1 strong acid

if $ka = 1 \rightarrow 10^{-3}$ moderate acid

if ka < 1 weak acid

pka:-

pka is defined as (–) negative Log of Ka.

$$pka = -\text{Log } ka$$

Smaller is the value of pka, larger will be ka and stronger would be the acid.

(b) **Levelling Effect:-**

If we use base to check strength of different acids but the base shows that all acids are of same strength, then the inability of base is called levelling effect. the base shows the strength of all acids at equal level.

Consider water act as a base, then it shows acidic strength as follows:

$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$

$$HI + H_2O \longrightarrow H_3O^+ + I^-$$

$$HBr + H_2O \longrightarrow H_3O^+ + Br^-$$

$$H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO_4^-$$

As all the acids have donated one H^+, the conjugate acids thus formed are all H_3O^+ which shows that the strength of all acids is same. This is levelling effect of water.

Q. || (a)

Hydrolysis:-

Hydrolysis is defined as "break down of water (H_2O) molecule into H^+ and OH^- ions."

(i) K^+ is derived from relatively strong base.

KOH is comparatively stronger base. Therefor it will not hydrolyze.

(ii) **NH_4^+**: conjugate acid of weak base

$$NH_4^+ + H_2O \longrightarrow NH_4OH + H^+ \quad \text{(Acidic solution)}$$

(iii) **CN^-**: conjugate of weak acid

$$CN^- + H_2O \longrightarrow HCN + OH^- \quad \text{(Basic solution)}$$

(b)

Extra Question

Unique property of buffer Solution:-

"A buffer Solution resists change in its pH even if small amount of strong acid or a base is added to it."

Types of acids dissolve in buffer Solution:-

A buffer Solution contains weak dissociating acid and the salt of that acid with a strong base.

	Name of acids:	Their Salts:
1	H_2CO_3	$H_2CO_3 + NaHCO_3$
2	H_3PO_4	$H_3PO_4 + NaH_2PO_4$
3	CH_3COOH	$CH_3COOH + CH_3COONa$

END.