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Chapter 9

ACIDS-BASES CHEMISTRY

ACIDS

1. It produces H^+ ions when dissolved in H_2O. (Arrhenius concept)
2. It has sour taste.
3. It turns blue litmus paper to Red.
4. They are classified into two types.

(a) Mineral Acid
→ stronger acid
→ HCl ; H_2SO_4 ; HNO_3

(b) Organic Acid
→ Weak Acid
→ Lactic acid, Tartaric acid, Maleic acid.

BASES

1. It produces OH^- ions when dissolved in H_2O. (Arrhenius concept)
2. It has bitter taste.
3. It turns red litmus paper to Blue.
4. They are classified into two types.

(a) Strong bases
[Group IA bases are stronger than Group II A bases]
→ $NaOH$; KOH ; $Ca(OH)_2$

(b) Weak bases
→ NH_4OH ; $Zn(OH)_2$; $Al(OH)_3$

5. Bases which are water soluble are called Alkalis.

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Q. All Alkalis are bases but all bases are not alkalis, why?

A. All alkalis are bases, but not all bases are alkalis. The key difference is their **solubility** and **strength**.

- Alkalis are highly soluble and strongly basic.
- Bases may not dissolve well in water and can be weak or strong.

∴ Alkalis can be considered as subset of bases.

$NaOH$: Alkali ey base

KOH : Alkali ey base

NH_3 : Base but not Alkali

$Fe(OH)_3$: Base but not alkali

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SALTS

The Substances obtained by **neutralization** of Acid ey base are called salts.

Acid + base ⇌ Salt + water

$HCl + NaOH \rightleftharpoons NaCl + H_2O$

TYPES OF SALTS:-

① Neutral salt ② Acidic salt ③ Basic salt

▷ **Neutral Salts** are obtained by reaction of Strong acid ey strong base.

$HCl + NaOH \rightleftharpoons NaCl + H_2O$

Strong acid Strong base neutral salt water

→ complete neutralization / replacement of H^+ & OH^-.

▷ **Acidic Salts** are obtained by reaction of Strong Acid ey weak base.

$HCl + NH_4OH \rightleftharpoons NH_4Cl + H_2O$

Strong acid weak base Acidic salt water

→ Incomplete neutralization / replacement of H^+.

▷ **Basic Salts** are obtained by reaction of weak acid and strong bases.

$2NaOH + H_2CO_3 \rightleftharpoons Na_2CO_3 + 2H_2O$

Strong base weak acid Acidic salts water

→ Incomplete neutralization / replacement of OH^-.

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ACIDIC, BASIC, & AMPHOTERIC SUBSTANCES

▷ **ACIDIC SUBSTANCE**, when dissolved in water produce Acid.

=> Non-metallic oxides are acidic substances.

CO_2 ; P_2O_5 ; N_2O_5 ; SO_2 ; SO_3

(i) $CO_2 + H_2O \longrightarrow H_2CO_3$ (carbonic acid)

(ii) $P_2O_5 + H_2O \longrightarrow H_3PO_4$ (phosphoric acid)

(iii) $N_2O_5 + H_2O \longrightarrow HNO_3$ (Nitric acid)

(iv) $SO_3 + H_2O \longrightarrow H_2SO_4$ (Sulphuric acid)

(v) $SO_2 + H_2O \longrightarrow H_2SO_3$ (Sulphurous acid)

∴ -ic → high oxidation state

∴ -ous → low oxidation state

▷ **BASIC SUBSTANCE**, when dissolved in water produce base.

=> Metallic oxides are basic substances.

Na_2O ; CaO ; MgO ; HgO

(i) $Na_2O + H_2O \longrightarrow NaOH$

(ii) $CaO + H_2O \longrightarrow Ca(OH)_2$

(iii) $MgO + H_2O \longrightarrow Mg(OH)_2$

(iv) $HgO + H_2O \longrightarrow Hg(OH)_2$

▷ **AMPHOTERIC SUBSTANCES**, react with acid as well as base. [it can act as acid as well as base : dual nature]

=> less electropositive Elements.

Al_2O_3 ; BeO ; ZnO ; PbO ; Cr_2O_3

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(i) Al_2O_3 + HCl → $AlCl_3$ + H_2O

base — acid — salt

Al_2O_3 + $NaOH$ → $NaAlO_2$ + H_2O

acid — base — salt

dual nature of amphoteric substances

(ii) Cr_2O_3 + HNO_3 → $Cr(NO_3)_3$ + H_2O

base — Acid — salt

Cr_2O_3 + KOH → K_2CrO_3 + H_2O

acid — base — salt

Differentiate between Acids & bases.

	ACIDS	BASES
1.	**Definition**	
	Donate H^+ ions (proton) or increase H^+ ion concentration.	Accept H^+ ions or increase OH^- ion concentration.
2.	**Chemical Properties**	
	Corrosive, sour, Good conductor.	Corrosive, bitter, Good conductor.
3.	**Litmus Paper**	
	Turns blue litmus into red.	Turns red litmus into blue.
4.	**pH**	
	pH is less than 7.	pH is greater than 7.
5.	**Ionization**	
	Release H^+ ion in solution.	Release OH^- ions in solution.
6.	**Reaction with metals**	
	React with metals to produce hydrogen gas.	Do not react with metals.

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7.	**Biological Role**	
	Digestion, protien breakdown	Neutralize stomach acid, maintain pH balance.
8.	**Environmental Impact**	
	Contribute to acid rain, harm aquatic life.	Help neutralize acid rain, maintain ecosystem balance.
9.	**Examples.**	
	Hydrochloric acid (HCl)	Sodium hydroxide ($NaOH$)
	Sulphuric acid (H_2SO_4)	Calcium hydroxide ($Ca(OH)_2$)
	Acetic acid (CH_3COOH)	Ammonia (NH_3)
	Citric acid ($C_6H_8O_7$)	Potassium hydroxide (KOH)

▷ **Differentiate between Organic & mineral acid.**

Organic Acid	**Mineral Acid**
1. **Sources**	
Derived from living organisms or plants	Derived from minerals or inorganic sources.
2. **Chemical Structure**	
Contain carbon & hydrogen atoms (CH_3COOH)	Contains hydrogen and a non-metal typically. (H_2SO_4)
3. **Properties**	
Generally weak, less corrosive, less toxic.	Strong, highly corrosive, and toxic.
4. **Solubility**	
Less soluble in water, more soluble in organic solvents.	highly soluble in water.

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5.	**pH**	
	Higher pH value (acetic acid : 2.4)	lower pH value. (sulphuric acid : 1)
6.	**Decomposition**	
	Decompose easily, forming carbon dioxide by water.	Do not decompose easily.
7.	**Reaction with bases**	
	Read slowly with bases.	React rapidly with bases.
8.	**Uses**	
	Food preservatives, cosmetics, pharmaceuticals.	Industrial processes, battery production, cleaning agents.
9.	**Examples**	
	Acetic acid (CH_3COOH)	Sulphuric acid (H_2SO_4)
	Citric acid ($C_6H_8O_7$)	Hydrochloric acid (HCl)
	Lactic acid ($C_3H_6O_3$)	Nitric acid (HNO_3)
	Tartaric acid ($C_4H_6O_6$)	Phosphoric acid (H_3PO_4)

BRONSTED & LOWERY CONCEPT

Acid: Proton donor

Base: Proton Acceptor

$$HCl + H_2O \longrightarrow H_3O^{\oplus} + Cl^{\ominus}$$

Acid Base Conjugate acid Conjugate base

Conjugate Base: When Acid donates a proton, it becomes conjugate base of that acid: -ive product

Cl^- in this example

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Conjugate Acid:- When a base accepts proton it becomes conjugate acid. : +ive product

H_3O^+ in this example

① $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Base acid conjugate acid conjugate base

② $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$

acid base conj. acid conj. base

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

acid base conj. base conj. acid

④ $H_2O + CO_3^{-2} \rightleftharpoons OH^- + HCO_3^-$

acid base conj. base conj. acid.

⑤ $HClO_4 + H_2O \rightleftharpoons ClO_4^- + H_3O^+$

Acid base conj. base conj. acid

▷ According to Bronsted - lowery concept:-

① Strong acid → weak conj base

weak acid → strong conj. base

② Strong base → weak conj. acid

weak base → strong conj acid

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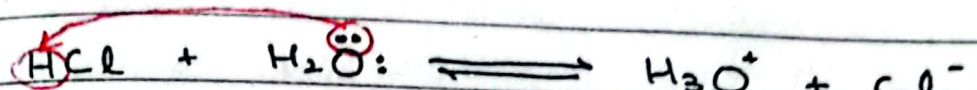
LEWIS CONCEPT

electron pair acceptor = Acid

electron pair donor = base



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

e⁻ pair e⁻ pair

acceptor: Acid donor: base

Acid base

① $Cl^- + AlCl_3 \longrightarrow [AlCl_4]^-$

base acid

② $H^+ + OH^- \longrightarrow H_2O$

acid base

③ $\ddot{N}H_3 + HCl \longrightarrow NH_4^+ + Cl^-$

base acid

④ $\ddot{N}H_3 + BF_3 \longrightarrow NH_3\text{-}BF_3$

base acid

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Q① Write limitations of :-

Arrhenius Theory :-

(i) It is restricted to aqueous solutions by does not apply to non-aqueous solvents or gaseous states.

(ii) It focuses only on H^+ by OH^- ions, ignoring other ions and molecular structures.

(iii) It fails to account for amphoteric substances,

(iv) It has limited predictive power regarding acid-base strength by behaviour.

② Bronsted-lowery Concept :-

(i) It only considers acid-base reactions involving H^+ ions.

(ii) It ignores electron pair interactions.

(iii) It is not effective in non-polar solvents or solid states.

(iv) It fails to explain acid-base behavior in multi-component systems.

③ Lewis concept :-

(i) It has overly broad definition of acids by bases, leading to ambiguity.

(ii) There is difficulty in predicting acid-base strength by reaction kinetics.

(iii) It ignores proton transfer reactions by aqueos solution behavior.

(iv) Limited aplicability to complex biological molecules by reaction mechanisms.

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IMPORTANT ACIDS AND BASES

Strong Acids

Hydrochloric acid	HCl
Nitric acid	HNO_3
Sulphuric acid	H_2SO_4
Perchloric acid	$HClO_4$

Weak Acids

Nitrous acid	HNO_2
Sulphurous acid	H_2SO_3
carbonic acid	H_2CO_3
phosphoric acid	H_3PO_4
propanoic acid	C_2H_5COOH

Strong Bases

Sodium Hydroxide	$NaOH$
Potassium Hydroxide	KOH
Lithium Hydroxide	$LiOH$

Weak Bases

Ammonium Hydroxide	NH_4OH
Calcium Hydroxide	$Ca(OH)_2$
Aluminium Hydroxide	$Al(OH)_3$

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STRENGTH OF ACID

The ability of a substance to donate / loose proton.

=> The lesser the time is taken to loose protons, the greater will be the degree of ionization.

$$HCl \rightleftharpoons H^+ + Cl^-$$

100% — 90% — 10%

High conc. of H^+ ions.

▷ Strong Acid:-

Acid which can donate proton to a higher degree than another acid.

Examples: HCl ; HNO_3 ; H_2SO_4 ; $HClO_4$

▷ Weak Acid:-

Acid which do not completely dissociate in aqueous solution & give lower concentration of H^+ ion are weak acid.

Examples:

HNO_2	nitrous Acid
H_2SO_3	sulphurous acid
H_2CO_3	carbonic acid
H_3PO_4	phosphoric acid
CH_3COOH	acetic acid
C_2H_5COOH	propanoic acid

-ous → Low oxidation state

-ic → High oxidation state

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STRENGTH OF BASES

The ability of a base to accept proton.

Strong Base:-

A base which can accept proton to a high degree than another base is relatively stronger base.

Example: NaOH ; KOH ; LiOH

(Group IA are more stronger than IIA)

Weak Base:-

A base which does not completely dissassociate in aqueous solution ey give lower concentration of OH^- ions are weak bases.

Example: NH_4OH Ammonium Hydroxide

$Ca(OH)_2$ Calcium Hydroxide

$Al(OH)_3$ Aluminium Hydroxide

(Group IIA is weaker than IA)

IONIZATION CONSTANT FOR WATER (K_w)

OR

IONIC PRODUCT OF WATER

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

$$K_c = \frac{[H_3O^+][OH^-]}{[H_2O][H_2O]} \quad \ldots (1)$$

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

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

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

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

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

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$[H]^+ = [OH]^- = 1 \times 10^{-7}$ M at 25°C for pure water

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

$K_w = 1 \times 10^{-14}$ ▷ at 25°C

K_w ↓ ionization constant of water

or

ionic product of water

$K_w \propto$ Temperature

small value = weak electrolyte Greater value = strong electrolyte

=> $K_w = 3.68 \times 10^{-4}$ at 40°C

As temperature is increased, K_w also increases that means it is a stronger electrolyte

$[H]^+ > [OH]^-$ → Acidic solution

$[H]^+ < [OH]^-$ → Basic solution

$[H]^+ = [OH]^-$ → Neutral

pH

Negative logarithm of H^+ ion concentration.

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

OR

Log of reciprocal of H^+ ion concentration.

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

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pH scale:-

neutral

0 — 1 — 2 — 3 — 4 — 5 — 6 — 7 — 8 — 9 — 10 — 11 — 12 — 13 — 14

↓ Strong acid | Acidic | ↓ weak acid | ↓ weak base | Basic | ↓ Strong base

Formulas:-

$pH + pOH = 14$ at 25°C

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

$1 \times 10^{-4} = [H]^+ [OH]^-$

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

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

Prove that / Derive $pH + pOH = 14$:-

ionization constant of water at 25°C

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

$[H]^+ [OH]^- = 1 \times 10^{-4}$ ∴ $K_w = 1 \times 10^{-14}$

Taking log on both sides

$\log [H]^+ + \log [OH]^- = -14 \log 10$ ∴ $\log a.b = \log a + \log b$

$\log [H]^+ + \log [OH]^- = -14 \times 1$ ∴ $\log x^n = n \log x$

Multiply (-1) on both sides

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

$pH + pOH = 14$ ∴ $pH = -\log [H]^+$

∴ $pOH = -\log [OH]^-$

Hence, proved!

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∴ Example # 9.2

$[OH]^- = 0.005$ M calculate $[H]^+ = ?$

$K_W = [H]^+ [OH]^-$

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

$= \frac{1 \times 10^{-14}}{0.005}$

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

∴ Example # 9.3

$pH = -\log 10^{-3}$

$pH = -(-3) \log 10$

$pH = 3$

∴ Example 9.4

NaOH = 0.062 M calculate pH

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

$= -\log [0.062]$

$pOH = 1.21$

$pH + pOH = 14$

$pH = 14 - pOH$

$= 14 - 1.21$

$pH = 12.79$

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Concept Assessment 9.2:-

Given:-

volume = 1 dm^3 mass of H_2SO_4 = 1.95 g

To find:-

pH = ?

Solution:-

$$\text{no. of moles} = \frac{\text{mass in g}}{\text{molar mass}}$$

$$= \frac{1.95}{98}$$

no. of moles (H_2SO_4) = 0.019 moles

$$\text{Molarity} = \frac{\text{no. of moles}}{\text{volume in } dm^3}$$

$$= \frac{0.019}{1}$$

Molarity = 0.019 M

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

$$= -\log [0.019]$$

pH = 1.4 **Answer**

TITRATION

Method to find volume of standard solution (Molarity known) react completely with known volume of another solution is called titration.

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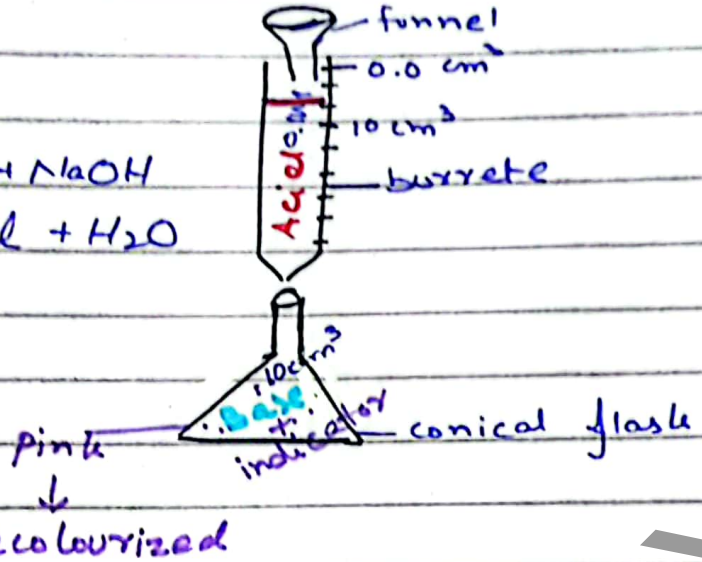

$HCl + NaOH \rightarrow NaCl + H_2O$

Concept assessment 9.3:-

25 cm^3 of 0.12 M NaOH neutralized with 30 cm^3 of HCl of unknown concentration. Calculate concentration & strength of HCl.

Data:- volume of NaOH = 25 cm^3 = V_2

molarity of NaOH = 0.12 M = M_2

volume of HCl = 30 cm^3 = V_1

To find:- concentration / Molarity of HCl = ? = M_1

strength of HCl = ?

Solution:- $HCl + NaOH \longrightarrow NaCl + H_2O$

$n_1 = 1$ $n_2 = 1$

HCl : NaOH

$$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2}$$

$$\frac{M_1 \times 30\ cm^3}{1} = \frac{0.12 \times 25\ cm^3}{1}$$

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$M_1 = \frac{0.12 \times 25\ cm^3}{30\ cm^3}$

$M_1 = 0.1\ M$

Strength = Molarity × molar mass

$= 0.1 \times 36.5\ g$

Strength = $3.65\ g/dm^3$

SELECTION OF INDICATORS

→ Indicators show whether the reaction is completed or not.

- When using strong Acid by strong base.

  Indicator : Phenolphthalein

  because the pH range /transition pH = 8 to 10

  (It changes colour between pH 8 to 10)

  colour of phenolphthalein in acids (1-7 pH) = Colourless

  colour of phenolphthalein in bases (8-14 pH) = pink

- When using weak Acid by strong base

  Indicator : Phenolphthalein

- When using Strong Acid by weak base.

  Indicator : Methyl Orange

  because the pH range /transition pH = 3.2 to 4.4

  colour of methyl orange in acids = Red

  colour of methyl orange in bases = Yellow

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• When using weak Acid by weak base.

Indicator : None

because there will be no ionization in solution and there will be no reaction between indicator and acid-base solution. And so, indicator will not work satisfactorily.

Strong by weak Acids or K_a

[ionization constant of acid (K_a)]

"Ability of acid to donate proton is called acidic strength"

→ More degree of ionization.

Strong acids: HCl ; H_2SO_4 ; $HClO_4$

Weak acids: H_3PO_4 ; CH_3COOH

acid

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

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

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

$\therefore K[H_2O] = K_a$

K_a ↓ ionization constant of water OR strength of acid

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

IMPORTANCE OF K_a :- K_a ∝ strength of acid

① if $K_a > 1$: acid is strong

② if $K_a = 1$ to 10^{-3} : moderately strong acid

③ if $K_a < 10^{-3}$: weak acid.

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For many acids, value of K_a is negative and it is difficult to write, so pK_a is used

$$pK_a = -\log K_a$$

or

$$pK_a = \log \frac{1}{K_a}$$

$$pK_a \propto \frac{1}{\text{strength of acid}}$$

eg: CH_3COOH (weak acid)

$K_a = 1.85 \times 10^{-5}$ $\quad$ $pK_a = 4.73$

→ pK_a value less ⟶ strong acid

→ pK_b value more ⟶ weak acid

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Strong & Weak Bases or K_b

[Ionization constant of Base (K_b)]

"Ability of base to accept proton is called basic strength"

→ how fastly / instantly base accepts proton.

→ high degree of ionization.

Example:

→ Alkali metals & Alkaline earth metals

▲ Strong bases : NaOH ; KOH

→ Weak bases : NH_4OH ; $Al(OH)_3$; $Zn(OH)_3$

K_b ∝ strength of base

↓

(Base dissasociation constant)

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

Base solvent

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

Water is solvent → constant

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

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

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To express K_b in convenient number, we take K_b in scale of **pK_b**

$$pK_b = -\log K_b$$

OR

$$pK_b = \log \frac{1}{K_b}$$

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

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

→ pK_b valueless ⟶ strong base

→ pK_b value greater ⟶ weak base

$$pK_b \propto \frac{1}{\text{strength of base } (K_b)}$$

Relationship between K_a & K_b

SLO Topic

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

② prove that $pK_a + pK_b = pK_w$

③ prove that $pK_a + pK_b = 14$

1: Proving that $K_a \propto \frac{1}{K_b}$

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

$$K_a = \frac{[H]^+[A]^-}{[HA]} \quad \text{(1)}$$

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

(conjugate base)

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

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

$\therefore K[H_2O] = K_b$ because we are ionizing base

$$K_b = \frac{[HA][OH^-]}{[A^-]} \quad \text{---------- (2)}$$

Multiplying eq (1) by (2)

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

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

$\therefore K_w = [H^+][OH^-]$

$$\boxed{K_a \times K_b = K_w}$$

$\therefore K_w$ = constant

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

$$\boxed{K_a \propto \frac{1}{K_b}}$$ Hence proved!

2: **Proving that $pK_a + pK_b = pK_w$**

$$K_a \times K_b = K_w$$

Applying log on both sides.

$\therefore \boxed{\log a \times b = \log a + \log b}$

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

Multiply -1 on both sides

$$-\log K_a + (-\log K_b) = -\log K_w$$

$$\boxed{pK_a + pK_b = pK_w}$$ Hence proved!

$\therefore pK_a = -\log[K_a]$

$\therefore pK_b = -\log K_b$

$\therefore pK_w = -\log K_w$

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3: Proving that $pK_a + pK_b = 14$:

$K_a \times K_b = K_w$

$K_a \times K_b = 1 \times 10^{-14}$ $\therefore K_w = 1 \times 10^{-14}$

∵ Taking log on both sides

$\log K_a + \log K_b = \log 1 + \log 10^{-14}$

$\log K_a + \log K_b = 0 + (-14 \log 10)$

$\log K_a + \log K_b = -14 \log 10$

$\log K_a + \log K_b = -14 (1)$

$\log K_a + \log K_b = -14$

∴ Multiply -1 on both sides

$-\log K_a + (-\log K_b) = -(-14)$ $\therefore pK_a = -\log K_a$

$\boxed{pK_a + pK_b = 14}$ $\therefore pK_b = -\log K_b$

Hence, proved! ☺

BUFFER SOLUTION

Solution in which there is no change in pH by adding small amount of acid or base or dilution or keeping it for a long time is called buffer solution.

Blood pH is 7.35 - 7.45

Stomach has 1 M HCl, but still, blood pH does not change. Blood is buffer solution.

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Importance of buffer solution:-

① Buffer solution is made to calibrate the pH meter.

② Biological catalysts i.e enzymes are pH dependent. if we are adding Acid/Base with enzyme it can be deactivated, then buffer is used.

TYPES OF BUFFER SOLUTION:-

① Acidic Buffer solution ② Basic Buffer solution

Acidic Buffer Solution:-

$$\underset{\text{(weak)}}{\text{Acid}} + \underset{\text{(strong)}}{\text{Base}} \longrightarrow \underset{\text{(strong base)}}{\text{Conjugate base}} + \underset{\text{(weak)}}{\text{Conjugate acid}}$$

Def:-

Mixture of weak acid by its salt [reacting with a strong base] is called acidic buffer solution.

① $\underset{\text{weak acid}}{CH_3COOH} + \underset{\text{strong base}}{NaOH} \longrightarrow \underset{\text{conj base}}{CH_3COONa} + \underset{\text{conj acid}}{H_2O}$

$[CH_3COOH \;\; CH_3COONa]$ → Acidic Buffer soln.

② $\underset{\text{weak acid}}{H_3PO_4} + 3\underset{\text{strong base}}{NaOH} \longrightarrow \underset{\text{conj base (salt)}}{Na_3PO_4} + 3H_2O$

$[H_3PO_4 \;\; Na_3PO_4]$ → Acidic buffer soln.

pH of acidic buffer is less than 7

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Basic Buffer Solution:-

It is mixture of weak base by its salt with strong acid.
[pH is greater than 7]

① $NH_4OH + NH_4Cl \longrightarrow$ Basic buffer

$HCl + NH_4OH \longrightarrow NH_4Cl + H_2O$
strong acid, weak base, salt

② $Al(OH)_3 + HCl \longrightarrow AlCl_3 + H_2O$
weak base, strong acid, salt

$[Al(OH)_3 / AlCl_3] \Rightarrow$ Basic Buffer

③ $Zn(OH)_2 + HCl \longrightarrow ZnCl_2 + H_2O$
weak base, strong acid, salt

$[Zn(OH)_2 / ZnCl_2] \Rightarrow$ Basic Buffer

BUFFER ACTION

SLO Topic

A method by which buffer can maintain its pH, is called buffer action.

⇒ Acidic Buffer

It is ionized as:

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$
move in solution; weak acid, weak electrolyte

$CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$
salt, strong electrolyte; move in solution

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▷ By adding acid in Acidic Buffer:-

H^+ from acid will combine with CH_3COO^- from salt and will form Acetic acid.

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

Acid — Acetate ion — Acetic acid

It is already present in buffer solution so pH will not change.

▷ By adding base in Acidic Buffer:-

OH^- from base will combine with H^+ from acidic buffer and form water. Water is weak electrolyte (neutral) so it will not further ionize and will not produce more H^+ or OH^-.

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

Base — Buffer solution — H_2O weak electrolyte (does not ionize)

No change in pH

⇒ Basic Buffer

It is ionized as:

$$\underset{\text{more in solution; weak base, weak electrolyte}}{NH_4OH} \rightleftharpoons NH_4^+ + OH^-$$

$$\underset{\text{salt, strong electrolyte}}{NH_4Cl} \rightleftharpoons \underset{\text{more in solution}}{NH_4^+ + Cl^-}$$

▷ By adding Acid in Basic Buffer:-

H^+ from acid will combine with OH^- from buffer and form water. It is a weak electrolyte (neutral) and will not ionize. pH will not be changed.

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

Acid — Buffer — water

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▷ By adding Base in Basic buffer:-

OH^- from base will combine with NH_4^+ from buffer by form NH_4OH. It is already present in solution so pH will not change.

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

Base Buffer Ammonium Hydroxide

APPLICATIONS OF BUFFER SOLUTION

1. Electroplating
2. Dyes preparation
3. Preparation of leather
4. Caliberation of pH meter
5. Bacteria Culture test

CALCULATION OF pH of BUFFER SOLUTION

$$pH = pKa + \log \frac{[salt]}{[acid]} \longrightarrow \text{Acidic Buffer}$$

$$pOH = pK_b + \log \frac{[salt]}{[base]} \longrightarrow \text{Basic Buffer}$$

▷ Henderson-Hesselbatch Equation

▷ **SELF CONCEPT 9.4:-**

Data:- $CH_3COOH = 0.09\ M$

$CH_3COONa = 0.11\ M$

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

To find:-

$pH = ?$

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Solution:- $pK_a = -\log K_a$

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

$\boxed{pK_a = 4.74}$

$\therefore pH = pK_a + \log \frac{[salt]\ (i.e.\ CH_3COONa)}{[Acid]\ (i.e.\ CH_3COOH)}$

$= 4.74 + \log \frac{(0.11)}{(0.09)}$

$\boxed{pH = 4.83}$ Answer

SALT HYDROLYSIS

Reaction of cations of Anions of salt with water is called salt hydrolysis.

OR

Breakdown of cations of anions in presence of water.

① eg Aqueous solution of NH_4Cl turns blue litmus to red.

$NH_4Cl + H_2O \rightleftharpoons NH_4OH + HCl$

salt — water — weak base — strong acid

↓ Acidic solution (Blue → Red)

SALT HYDROLYSIS IS REVERSE OF NEUTRALIZATION

② Aqueous solution of K_2CO_3 turns red litmus to blue

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

salt — water — strong base — weak acid

↓ basic solution (Red → Blue)

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③ Aqueous Solution of NaCl has no change on litmus paper.

$NaCl + H_2O \rightleftharpoons NaOH + HCl$

(NaOH: strong base, HCl: strong acid)

No change on litmus paper

TYPES OF SALT HYDROLYSIS

① **Neutral Salt:-**

Salt of strong acid and strong base. It does not hydrolyze.

Strong Acid + Strong base → Neutral salt

Examples:- NaCl ; Na_2SO_4 ; KNO_3

The pH of neutral salt is 7

② **Acidic Salt:-**

Salt of strong acid and weak base.

Strong Acid + Weak base → Acidic Salt

Examples:- $CuSO_4$; NH_4Cl ; NH_4NO_3

The pH of Acidic salt is less than 7

③ **Basic Salt:-**

Salt of strong base and weak acid.

Strong base + weak acid → Basic salt

Examples:- Na_2CO_3 ; CH_3COONa ; NaCN ; Na_2S

The pH of Basic salt is greater than 7

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④ Neutral / Acidic / Basic Salt:-

salt of weak acid and weak base can be neutral, acidic or basic.

Weak acid + Weak base ⟶ Salt [neutral / acidic / basic]

→ It depends on value of K_a & K_b

$K_a > K_b$ ⟶ Acidic salt

$K_a < K_b$ ⟶ Basic salt

$K_a \approx K_b$ ⟶ neutral

Example:- $\underset{\text{weak acid}}{H_2CO_3} + \underset{\text{weak base}}{NH_4OH} \longrightarrow (NH_4)_2CO_3 + H_2O$

SOLUBILITY PRODUCT (K_{sp})

• **Sparingly Soluble salts** (ionic compound): A salt that does not completely dissolve in water.

→ Excess of sparingly soluble salt is mixed with water. Some of it is dissolved and rest will settle at the bottom. Dynamic equilibrium is established between insoluble solid compound (settled at bottom) and its ions (soluble) in a saturated solution.

$$\underset{\text{(sparingly soluble)}}{CaF_{2(s)}} \rightleftharpoons \underset{\text{(soluble)}}{Ca^{+2}_{(aq)}} + 2F^-_{(aq)}$$

equilibrium is established

"Product of Molar concentration of opposite charged ions in equilibrium with its solid state is called solubility product" (K_{sp})

→ K_{sp} is only written for partially / sparingly soluble salt.

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→ K_{sp} is temperature dependent.

$$CaF_{2(s)} \rightleftharpoons Ca^{+2}_{(aq)} + 2F^-_{(aq)}$$

Equilibrium is established between solid form of salt and aqueous (soluble) form of salt.

K_c can be written as:-

$$K_c = \frac{[Ca^{+2}][F^-]^2}{[CaF_2]}$$

Since CaF_2 is slightly soluble, it's amount is constant.

$$K_c[CaF_2] = [Ca^{+2}][F^-]^2$$

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

General form:-

$$A_mB_m \rightleftharpoons mA^{+n} + nB^{-m}$$

$$\boxed{K_{sp} = [A^{+n}]^m[B^{-m}]^n}$$

Eg:- $PbCl_2 \rightleftharpoons Pb^{+2} + 2Cl^-$

$$\boxed{K_{sp} = [Pb^{+2}][Cl^-]^2}$$

Concept Assessment 9.5:-

(i) a) Iron (II) Hydroxide.

$$Fe(OH_2) \rightleftharpoons Fe^{+2} + 2OH^-$$

$$\boxed{K_{sp} = [Fe^{+2}][OH^-]^2}$$

b) Calcium Sulphate

$$CaSO_4 \rightleftharpoons Ca^{+2} + SO_4^{-2}$$

$$\boxed{K_{sp} = [Ca^{+2}][SO_4^{-2}]}$$

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(ii) $PbSO_4 \rightleftharpoons [Pb^{+2}][SO_4^{-2}]$

$K_{sp} = 1.98 \times 10^{-8}$ at 25 °C

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

$1.96 \times 10^{-8} = [Pb^{+2}][SO_4^{-2}]$

$1.96 \times 10^{-8} = x \cdot x$

$1.96 \times 10^{-8} = x^2$

Taking $\sqrt{}$ on b.s

$\sqrt{1.96 \times 10^{-8}} = \sqrt{x^2}$

$x = 1.4 \times 10^{-4}$ M

Solubility of $PbSO_4$ is 1.4×10^{-4} M.

(iii) Given:- Initial concentration of:

$[Ca^{+2}] = 1 \times 10^{-9}$ M

$[PO_4^{-3}] = 1 \times 10^{-9}$ M

$K_{sp} = 1.2 \times 10^{-29}$ at 25 °C

To find:- Will precipitation occur?

Solution:-

$Ca_3(PO_4)_2 \rightleftharpoons 3Ca^{+2} + 2PO_4^{-3}$

$Q' = [Ca^{+2}]^3 [PO_4^{-3}]^2$

$Q' = (1 \times 10^{-9})^3 (1 \times 10^{-9})^2$

$Q' = 1 \times 10^{-45}$

While $K_{sp} = 1.2 \times 10^{-29}$

As $Q' < K_{sp}$, so precipitation will not occur.

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PRECIPITATION REACTION

Reaction in which two or more than two solutions are mixed together to produce insoluble product.

eg:- $AgNO_{3(aq)} + NaCl_{(aq)} \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$

Solutions — insoluble in water [forms ppt] — soluble

- **CASE 1:-**

  (ion product) $Q' < K_{sp}$

  unsaturated solution → no ppt

- **CASE 2:-** $Q' = K_{sp}$

  saturated solution → no ppt

- **CASE 3:-** $Q' > K_{sp}$

  Supersaturated solution → ppt will form.

ion product, Q' : Molar concentration of opposite charged ions with solid state not in Eq. position

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COMMON ION EFFECT

Decrease in solubility or ionization of an electrolyte by the addition of another electrolyte having a common ion is called common ion effect.

Eg:- $NaCl \rightleftharpoons Na^+ + Cl^-$ (weak electrolyte)

$HCl \rightleftharpoons H^+ + Cl^-$ (strong electrolyte)

When concentration of substance increase on product side it will undergo reverse direction ey NaCl is precipitated out.

(undissolved) $NaCl \rightleftharpoons Na^+ + Cl^-$ ← left side; product ↑

(precipitates out) $HCl \rightleftharpoons H^+ + Cl^-$

- substances having more ionizing power are more sowble, whereas substances having less ionizing power are left undissowed.

1. (ppt) $NaF \rightleftharpoons Na^+ + F^-$ (Na salt)
   (dissolve) $HF \rightleftharpoons H^+ + F^-$ (Acid)

2. (ppt) $KClO_4 \rightleftharpoons K^+ + ClO_4^-$ (salt)
   (dissolve) $KCl \rightleftharpoons K^+ + Cl^-$ (acid)

   Ka ∝ Acidic strength

3. (ppt) $NH_4OH \rightleftharpoons NH_4^+ + OH^-$ (base)
   (dissolve) $NH_4Cl \rightleftharpoons NH_4^+ + Cl^-$ (salt)

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▷ SELF CONCEPT 9.6:-

1. $NH_4OH \rightleftharpoons NH_4^+ + OH^-$

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

→ NH_4^+ is common ion. It's concentration increases so equilibrium will shift towards left.

→ NH_4Cl is strong electrolyte, so it will dissolve.

→ NH_4OH is weak electrolyte and form precipitates.

2. $H_2CO_3 \rightleftharpoons 2H^+ + CO_3^{-2}$

   $Na_2CO_3 \rightleftharpoons 2Na^+ + CO_3^{-2}$

→ CO_3^{-2} is common ion. Its concentration increases so equilibrium will shift towards left.

→ Na_2CO_3 is strong electrolyte, so it will dissolve.

→ H_2CO_3 is weak electrolyte, so it will form precipitates / crystals.