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* Acid-Base Concept:

(1) Arrhenius concept
(2) Lewis concept
(3) Bronsted - Lowry

① Arrhenius Concept →

Acc. to Arrhenius,

Acid → is the substance which can be furnished into H^+/proton & corresponding anion in aqueous medium.

Eg → HCl, H_2SO_4, H_2CO_3, HBr, HI, HF, HPO_4, HNO_3, H_2SO_3, $H_2S_2O_7$, $H_2S_2O_8$, $HClO_4$, $HBrO_4$, $HOCl$, HOF, $B(OH)_3$ etc.

A substance which has ionisable H are said to be Arrhenius Acid i.e. all protic acids are Arrhenius Acid.

Base → Substance which can be furnished into OH^- (Hydroxyl ion) & corresponding ion in aqueous medium.

Eg → $LiOH$, KOH, $RbOH$, $CsOH$, $NaOH$, $Be(OH)_2$, $Ca(OH)_2$, $Mg(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$, $Fe(OH)_2$, $NH_4(OH)_2$, $Al(OH)_3$, $Cr(OH)_3$

② Lewis Concept →

Acid → These are e⁻ deficient species which can accept lone pair of e⁻

All e⁻ deficient species are Lewis Acid

Eg → All Cations

→ Compounds having incomplete octet of central atom, Eg → BF_3, BCl_3, BI_3, BBr_3, BH_3, BCl_2, BH_2 etc.

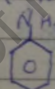
Base → Substance which can donate its l.p of e⁻

Electron rich comp.

Eg → All Anions

→ l.p. containing compounds

Eg – $\ddot{N}H_3$, $R-\ddot{N}H_2$, $R_2\ddot{N}H$, $R_3\ddot{N}$, $R-\ddot{O}-R$, $R-\ddot{O}H$



③ Bronsted-Lowry Concept →

Proton donor species are Acids.

Ex: $B(OH)_3$

It is monovalent acid

Substance which can accept H^+ ion is Base.

Eg→ $\ddot{N}H_3 + H^+ \longrightarrow NH_4^+$

NH_4 is Lewis base as well as Bronsted Base.

** → $H_2O \rightleftharpoons H^+ + OH^-$ (Acid)

$H_2O + H^+ \rightleftharpoons H_3O^+$ (Base)

→ $H_2S \rightleftharpoons H^+ + HS^-$ (Base)

$HS^- \rightleftharpoons H^+ + S^{-2}$ (Acidic)

| $AH \rightleftharpoons H^+ + A^-$ |

A^- is conjugate base of AH
AH is —"— acid of A^-

| Conjugate acid $\underset{H^+ \text{ gain}}{\overset{H^+ \text{ loss}}{\rightleftharpoons}}$ Conjugate Base |

→ HS^- is Conjugate base of H_2S
HS^- —"— acid of & S^{-2}
H_2S —— acid of HS^-
S^{-2} —"— base of HS^-

→ Bronsted Concept is based on Relative Strength of given substance.

→ Bronsted concept needs no solvent to explain acidic & basic strength.

→ A substance may have acid as well as base depends on its relative strength wrt to which ~~it is~~ its acidic strength compared.

→ Conjugate acid & base are differ by a single proton to each other.

$$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$$
I → II

$$H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{-2}$$
III

$$HPO_4^{-2} \rightleftharpoons H^+ + PO_4^{-3}$$
IV

I & II are conjugated A-B to each oth.
II & III —— " —— "
III & IV —— " —— "

Eg ⇒ $CH_4 < NH_3 < H_2O < HF$
↓ ↓ ↓ ↓
$CH_3^{\ominus} < NH_2^{\ominus} < OH^- < F^-$
EN ↑
Stablity of ⊖ ↑

⊖ ⇒ money
⊕ ⇒ loan
High E.N. ⇒ Miser / कंजूस
Less E.N. ⇒ दिलदार

Stablity of -ve charge $\propto$ EN
—"— +ve charge $\propto \frac{1}{EN}$
—"— of charge $\propto$ Size
Reactivity of Reactant $\propto$ Stablity of product
Acidic nature of an acid $\propto$ Stablity of its conjugate base.

$HF < HCl < HBr < HI$
↓ ↓ ↓ ↓
F^- Cl^- Br^- I^-
Size ↑
Stablity of Charge ↑

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* Ostwald's Dilution Theory:

$$AB \longrightarrow A^+ + B^-$$

$\alpha = 100\%$, strong electrolyte

$$AX \rightleftharpoons A^+ + X^-$$

weak

On Dilution,

V↑, Eq^m shifts to more no. of mole
⇒ —, — forward
⇒ Dissociation of reactant ↑

→ Acc. to this Theory, dissociation of a weaker electrolyte is directely proport. to it dissociation
→ Ostwald's dilution theory applicable only for weaker electrolytes.

• App. of Ostw. Dil. Theory:

→ Diss. of weaker acid:

	HA	$\rightleftharpoons$	H^+	$+$	A^-
Initial Conc. (M)	C		0		0
α^{HA}					
At eq^m	$C - C\alpha$		$C\alpha$		$C\alpha$

$$K_{eq} = K_a = \frac{[H^+][A^-]}{[HA]} = \frac{C\alpha \times C\alpha}{C - C\alpha} = \frac{C^2\alpha^2}{C(1-\alpha)}$$

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$$K_a = \frac{C\alpha^2}{1-\alpha}$$

$\alpha << 1$ for weaker acid
$1-\alpha \approx 1$

$$K_a = C\alpha^2$$

$$\alpha = \sqrt{\frac{K_a}{C}}$$

$$[H^+] = C\alpha = C\alpha\sqrt{\frac{K_a}{C}} = \sqrt{K_a \times C}$$

$$pH = -\log(H^+)$$
$$= -\log[K_a \times C]^{\frac{1}{2}}$$
$$= \frac{1}{2}\{-\log(K_a \times C)\}$$
$$= \frac{1}{2}\{-\log K_a - \log C\}$$
$$= \frac{1}{2}[pK_a - \log C]$$

$$pH = \frac{1}{2}pK_a - \frac{1}{2}\log C$$

Ex→

$$BOH \rightleftharpoons B^+ + OH^-$$

Initial conc. (M) $\quad C \quad 0 \quad 0$

α^{BOH}

At Eq^m $\quad C - C\alpha \quad C\alpha \quad C\alpha$

$$K_{eq} = K_b = \frac{[B^+][OH^-]}{[BOH]} = \frac{C\alpha \times C\alpha}{C - C\alpha} = \frac{C^2\alpha^2}{C(1-\alpha)}$$

$$\left\{K_b = \frac{C\alpha^2}{1-\alpha}\right\}$$

$\alpha << 1$ for weaker base
$1-\alpha \approx 1$

$$\alpha = \sqrt{\frac{K_b}{C}}$$

$$[OH^-] = C\alpha = C \times \sqrt{\frac{K_b}{C}} = \sqrt{K_b \times C}$$

$$p^{OH} = -\log [OH^-]$$
$$= -\log [K_b \times C]^{\frac{1}{2}}$$
$$= \frac{1}{2} [-\log K_b - \log c]$$

$$\boxed{p^{OH} = \frac{1}{2} p^{K_b} - \frac{1}{2} \log c}$$

- Acidic strength of weaker acid $\propto [H^+]$

$(HA)_1$, $[H^+]_1 = \sqrt{K_{a_1} \times C_1}$

$(HA)_2$, $[H^+]_2 = \sqrt{K_{a_2} \times C_2}$

$$\frac{[H^+]_1}{[H^+]_2} = \frac{\sqrt{K_{a_1} \times C_1}}{\sqrt{K_{a_2} \times C_2}}$$

① if both acidic solⁿ are isomolar

$C_1 = C_2$

$$\boxed{\text{Acidic strength} \propto K_a}$$

② An acid having diff. conc.

Acidic strength $\propto$ Conc.

③ Isohydric Solⁿ

Solⁿ having equal conc. of $[H^+]$

$$[H^+]_1 = [H^+]_2$$

$$\sqrt{K_{a_1} \times C_1} = \sqrt{K_{a_2} \times C_2}$$

$$\boxed{K_{a_1} C_1 = K_{a_2} C_2}$$

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Q. Calculate pH & α of 10^{-3} M HCN ($K_a = 10^{-5}$)

Monovalent acid

$\Rightarrow$ $K_a = C\alpha^2$

$10^{-5} = 10^{-3} \times \alpha^2$

$\alpha = \sqrt{\frac{10^{-5}}{10^{-3}}} = 10^{-1}$

$\alpha = 0.1 = 10\%$

$pH = -\log(H^+)$

$= -\log(10^{-4})$

$= 4$

$(H^+) = C\alpha$

$= 10^{-3} \times 10^{-1}$

$= 10^{-4}$

Q. Calculate pH of 10^{-2} M H_2CO_3 which dissociates 25% upto eq^m.

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

$C \quad\quad 0 \quad\quad 0$

$C - C\alpha \quad\quad 2C\alpha \quad\quad C\alpha$

$[H^+] = 2C\alpha$

$= 2 \times 10^{-2} \times 0.25$

$= 0.5 \times 10^{-2}$

$= 5 \times 10^{-3}$

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

$= -\log(5 \times 10^{-3})$

$pH = -[\log 5 + \log 10^{-3}]$

$= 3 - \log 5$

$= 3 - 0.7$

$= 2.3$

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Q: Calculate pH of 10^{-3} M H_3PO_4 if $\alpha_1 = 100\%$, $\alpha_2 = 50\%$, $\alpha_3 = 0\%$.

→ $H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$

C	0	0
$C - C\alpha_1$	$C\alpha_1$	$C\alpha_1$

$H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{-2}$

$C\alpha_1$	0	0
$C\alpha_1 - C\alpha_1\alpha_2$	$C\alpha_1\alpha_2$	$C\alpha_1\alpha_2$

$HPO_4^{-2} \rightleftharpoons H^+ + PO_4^{-3}$

$C\alpha_1\alpha_2$	0	0
$C\alpha_1\alpha_2 - C\alpha_1\alpha_2\alpha_3$	$C\alpha_1\alpha_2\alpha_3$	$C\alpha_1\alpha_2\alpha_3$

$[H^+] = C\alpha_1 + C\alpha_1\alpha_2 + C\alpha_1\alpha_2\alpha_3$
$= C\alpha_1[1 + \alpha_2 + \alpha_2\alpha_3]$
$= C \times 1\,[1 + 0.5 + 0.5 \times 0]$
$= C\,[1.5]$

$[H^+] = 1.5C$
$= 1.5 \times 10^{-3} M$

$pH = -\log(1.5 \times 10^{-3}) \Rightarrow = -\log\left[\frac{3}{2} \times 10^{-3}\right]$
$= 3 + \log 2 - \log 3$
$= 3 + 0.3010 - 0.4770$
$= 2.83$

$pH = -\log(H^+)$
$= -\log\left[\frac{a}{b} \times 10^{-c}\right]$

$$pH = c + \log b - \log a$$

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Q. Calculate degree of dissociation of $10^{-3}M$ acetic acid in a sol^n having pH = 4.

→ $pH = -\log(H^+) = 4$

$\log_{10}(H^+) = -4$

$(H^+) = 10^{-4}$

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

C   0   0

$C - C\alpha$ $C\alpha$ $C\alpha$

$[H^+] = C\alpha$

$10^{-4} = 10^{-3} \times \alpha$

$\alpha = 10^{-1} = 10\%$

Q. Calculate pH of:

(i) 10^{-8} M NaOH   (ii) 10^{-8} M HCl.

→ (i) 10^{-8} M NaOH

$[OH^-]_{Total} = [OH^-]_{NaOH} + [OH]_{H_2O}$ | $H_2O \rightleftharpoons H^+ + OH^-$ (10^{-7}, 10^{-7})

$= 10^{-8} + 10^{-7}$

$= 10^{-8}(1 + 10)$

$= 11 \times 10^{-8}$

$[OH^-] = 1.1 \times 10^{-7}$

$pOH = -\log(OH^-)$

$= -\log(1.1 \times 10^{-7})$

$= 7 - \log 1.1$

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$pOH = 7 - 0.03 = 6.97$

$pH = 14 - pOH$
$= 14 - 6.97$
$= 7.03$

(iii) $[H^+]_{total} = [H^+]_{HCl} + (H^+)_{H_2O}$
$= 10^{-8} + 10^{-7}$
$= 10^{-8}(1 + 10)$
$= 10^{-8} \times 11$
$[H^+] = 1.1 \times 10^{-7}$
$pH = -\log[H^+]$
$= -\log(1.1 \times 10^{-7})$
$= 7 - \log 1.1$

$$pH = 7 - 0.03 = 6.97$$

* WATER:-
Very weak electrolyte
At room temp.,
(1) Conc. of H_2O / molarity of H_2O = 55.5 M
H_2O molecules in 1 litre of $H_2O = 55.5 \times N_A$
H^+ ions in 1 litre of $H_2O = 10^{-7} \times N_A$
$[H_2O] : [H^+]$
$55.5 \times N_A = 10^{-7} \times N_A$

$$55.5 \times 10^{-7} : 1$$

$\alpha_{H_2O} = \frac{\text{Dissociated molecules}}{\text{Initial Molecules}} = \frac{1}{55.5 \times 10^{+7}} = \frac{18}{1000 \times 10^{7}} = 1.8 \times 10^{-9} = 1.8 \times 10^{-7}\%$

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② It is neutral

$(H^+) = (OH^-)$

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

$$K = \frac{[H^+][OH^-]}{[H_2O]} = \frac{10^{-7} \times 10^{-7}}{55.5}$$

$$= \frac{10^{-14} \times 18}{1000}$$

$\boxed{K = 1.8 \times 10^{-16}}$ Dissociation const. of H_2O

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

$K[H_2O] = [H^+][OH^-]$ = Ionic Product of $H_2O = K_\omega$

$K_\omega = [H^+][OH^-]$ $\quad [\,[H^+] = [OH^-]\,]$

$= [H^+]^2$

$= [OH^-]^2$

$= [10^{-7}]^2$

$\boxed{K_\omega = 10^{-14}}$

$K_\omega = [H^+][OH^-]$

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

$pK_\omega = 2pH = 2pOH$

$pH = pOH = \frac{pK_\omega}{2} = \frac{14}{2} = 7$

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

$\therefore pH = pOH$

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• Effect of Temprature on dissociation of H_2O

T↑ Dissociation of H_2O ↑

$\alpha\uparrow$, $[H^+]\uparrow$, $[OH^-]\uparrow$, $K_w\uparrow$

$p^{K_w}\downarrow$, $pH\downarrow$, $p^{OH}\downarrow$

Q. At 90°C ionic product of water is $10^{-13.62}$. Calculate pH of water & identify its nature.

→ Neutral

~~pH 6.81~~

$$pH = \frac{10^{-13.62}}{2} = \frac{13.62}{2} = 6.81$$

★ Salt Hydrolysis:

Salt = Cation + Anion

Basic Radical ← [+ | −] → Acidic Radical

Neutralisation ⇒ Reaction b/w acid & base to form corresponding Salt & water.

Ex:

$$HCl + KOH \underset{\text{backward}}{\overset{\text{forward}}{\rightleftharpoons}} KCl + H_2O$$

forward ← Neutralisation

backward ← Hydrolysis

Hydrolysis ⇒ Ionisation of salt in presence of H_2O

⇒ It is inverse process of neutralisation

$$K_h = \frac{1}{K_N}$$

- Strong Acid → HCl, HBr, HI, H_2SO_4, HNO_3, Picric acid, $HClO_4$, $Ca(OH)_2$
- Strong Base → NaOH, KOH, RbOH, CsOH, $Ba(OH)_2$, $Sr(OH)_2$
- Types of Salts :

① SASB (St. Acid St. Base)

$CaCl_2$, NaCl, Na_2SO_4, KNO_3

② SAWB (St. Acid Wk. Base)

LiCl, $MgSO_4$, NH_4NO_3

③ WASB (Wk. Acid St. Base)

KCN, CH_3COONa, $\begin{matrix} COOK \\ | \\ COOK \end{matrix}$, Na_2CO_3

④ WAWB (Wk. Acid Wk. Base)

NH_4CN, $MgCO_3$, $\begin{matrix} COO^- \\ | \\ COO^- \end{matrix} Mg^{+2}$

Strong → Nature
Weak → Hydrolysis

SA → $p^H < 7$, acidic
SB → $p^H > 7$, basic
WB → Cation
→ cationic Hydrolysis
WA → anion
→ anionic Hydrolysis

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WAWB (Wk. Acid Wk. Base)

(i) $K_a = K_b$, Neutral Hydrolysis
(ii) $K_a > K_b$, Slightly acid
Cationic-anionic Hydrolysis $pH \Rightarrow 5$ to 7
(iii) $K_a < K_b$, Slightly basic ,
Anionic-Cationic Hydrolysis $pH \Rightarrow 7$ to 9

SASB type Salts (St. Acid St. Base):

$$NaCl + H_2O \rightleftharpoons NaOH + HCl$$

Strong electrolytes

$$Na^+_{aq} + Cl^-_{aq} + H_2O \rightleftharpoons Na^+_{aq} + OH^- + H^+ + Cl^-$$

Spectator ions

Note – During Hydrolysis of salt pH must differ from 7, i.e. there should be change in conc. of H^+ or OH^- ions or both.
SASB type salt donot exhibit Hydrolysis. They simply dissolve in water.

WASB type Salts:

KCN

$$KCN + H_2O \rightleftharpoons KOH + HCN$$

Strong; KOH strong; HCN weaker

$$K^+_{aq} + CN^-_{aq} + H_2O \rightleftharpoons K^+_{aq} + OH^- + HCN$$

$$CN^-_{aq} + H_2O \rightleftharpoons OH^- + HCN \quad \text{Basic}$$

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- h = Degree of Hydrolysis
- K_h = Hydrolysis const.
- K_a = Dissociation const. of weak acid
- K_B = Diss. const. of weak base
- K_w = Ionic product of H_2O

	$CN^-_{(aq)}$	+ H_2O	⇌	HCN	+ OH^-
Initial Conc.	C	(excess)		0	0
At Eq^m	$C - Ch$			Ch	Ch

$$K_h = \frac{[HCN][OH^-]}{[CN^-]} = \frac{Ch \times Ch}{C - Ch} = Ch^2 \quad -(1)$$

$$K_h = Ch^2$$

$$\boxed{h = \sqrt{\frac{K_h}{C}}} \quad -(2)$$

$$K_w = [H^+][OH^-] \quad -(3)$$

Weaker dissociates

$$HCN \rightleftharpoons H^+ + CN^-$$

$$K_a = \frac{[H^+][CN^-]}{[HCN]} \quad -(4)$$

Eq (3) = Eq (1) × Eq (4)

$$\boxed{K_w = K_h \times K_a} \quad -(5)$$

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$$K_h = \frac{K_w}{K_a} = Ch^2$$

$$h = \sqrt{\frac{K_w}{K_a \times C}} \quad \text{— (6)}$$

$[OH^-] = Ch$

$$[OH^-] = \sqrt{\frac{K_w \times C}{K_a}} \quad \text{— (7)}$$

$[H^+] = \frac{K_w}{[OH^-]}$ from Eq (3)

$$[H^+] = \sqrt{\frac{K_w \times K_a}{C}} \quad \text{— (8)}$$

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

$pH = \frac{pK_w}{2} + \frac{1}{2}pK_a + \frac{1}{2}\log C$

$$pH = 7 + \frac{1}{2}pK_a + \frac{1}{2}\log C \quad \text{— (9)}$$

- ~~WAW~~ SA WB:

NH_4Cl

$NH_4Cl + H_2O \rightleftharpoons NH_4OH + HCl$

Strong electrolyte — weaker — strong

$NH_4^+{}_{aq} + Cl^-{}_{aq} + H_2O \rightleftharpoons NH_4OH + H^+ + Cl^-{}_{aq}$

	NH_4^+	$+ H_2O$	$\rightleftharpoons$	NH_4OH	$+ H^+$
Initial conc. —	C	(excess)		0	0
At Eqm —	C − Ch			Ch	Ch

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$$K_h = \frac{[NH_4OH][H^+]}{[NH_4^+]} \quad \text{—— (1)}$$

$$K_h = Ch^2$$

$$h = \sqrt{\frac{K_h}{C}}$$

$$K_w = (H^+)(OH^-) \quad \text{— (3)}$$

Weaker base dissociates

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

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]} \quad \text{— (4)}$$

Eq (3) = Eq (1) × Eq (4)

$$K_w = K_h \times K_b \quad \text{— (5)}$$

$$K_h = \frac{K_w}{K_b} = Ch^2 \quad \text{— (6)}$$

$$h = \sqrt{\frac{K_w}{K_b \times C}} \quad \text{— (7)}$$

$$[H^+] = Ch$$

$$[H^+] = \sqrt{\frac{K_w \times C}{K_b}} \quad \text{— (8)}$$

$$pH = 7 - \frac{1}{2} p^{K_b} - \frac{1}{2} \log C \quad \text{— (9)}$$

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· **WAWB type of salts** →

NH_4CN

$NH_4CN + H_2O \rightleftharpoons NH_4OH + HCN$
Strong

$NH_4^+{}_{aq} + CN^-{}_{aq} + H_2O \rightleftharpoons NH_4OH + HCN$

Initial conc.	C	C	(excess)	0	0
At Eq^m	$C - Ch$	$C - Ch$		Ch	Ch

$$K_h = \frac{[NH_4OH][HCN]}{[NH_4^+][CN^-]} \quad \text{— ①}$$

$$= \frac{Ch \times Ch}{C-Ch \times C-Ch} = \frac{C^2h^2}{C(1-h)C(1-h)}$$

$K_h = h^2$

$h = \sqrt{K_h}$ — ②

$K_w = [H^+][OH^-]$ — ③

Weaker acid

$HCN \rightleftharpoons H^+ + CN^-$

$$K_a = \frac{[H^+][CN^-]}{[HCN]} \quad \text{— ④}$$

Weak base

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

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$$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]} \quad \text{— (5)}$$

Eq (3) = Eq (1) + Eq (4) + Eq (5)

$K_w = K_h \times K_a \times K_b$

$$K_h = \frac{K_w}{K_a \times K_b} \quad \text{— (6)}$$

$$h = \sqrt{\frac{K_w}{K_a \times K_b}} \quad \text{— (7)}$$

from eq. (4),

$$\frac{K_a \times [HCN]}{[CN^-]} = [H^+]$$

$$\frac{K_a \times Ch}{C - Ch} = [H^+]$$

$$K_a \times h = [H^+] \quad \text{— (8)}$$

$[H^+] = K_a \times h$

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

$$[H^+] = \sqrt{\frac{K_w \times K_a}{K_b}} \quad \text{— (9)}$$

$$pH = 7 + \frac{1}{2} pK_a - \frac{1}{2} pK_b$$

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Types of Salts	K_h	h	$[H^+]$	pH
(1) SASB	— No comments			$pH = 7$
(2) SAWB	$K_h = Ch^2 = \frac{K_w}{K_b}$	$h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{K_w}{K_b \times C}}$	$[H^+] = \sqrt{\frac{K_w \times C}{K_b}}$	$pH = 7 - \frac{1}{2} pK_b - \frac{1}{2} \log C$
(3) WASB	$K_h = Ch^2 = \frac{K_w}{K_a}$	$h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{K_h}{K_a \times C}}$	$[H^+] = \sqrt{\frac{K_w \times K_a}{C}}$	$pH = 7 + \frac{1}{2} pK_a + \frac{1}{2} \log C$
(4) WAWB	$K_h = h^2 = \frac{K_w}{K_a \times K_b}$	$h = \sqrt{K_h} = \sqrt{\frac{K_w}{K_a \times K_b}}$	$[H^+] = \sqrt{\frac{K_w \times K_a}{K_b}}$	$pH = 7 + \frac{1}{2} pK_a - \frac{1}{2} pK_b$

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for WAWB

① If, $K_a = K_b$

⇒ Neutral Hydrolysis

⇒ pH = 7

② If, $K_a > K_b$

⇒ Cationic-anionic Hydrolysis

⇒ Slightly less than 7

③ If $K_a < K_b$

⇒ Anionic Cationic Hydrolysis

⇒ Slightly more than 7.

Q. Calculate K_h, h & pH of 10^{-3} M KCN solⁿ if K_a for HCN is 10^{-6}.

→ KCN

WASB,

$$K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{10^{-6}} = 10^{-8}$$

$$K_h = Ch^2$$

$$h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{10^{-8}}{10^{-3}}} = \sqrt{10^{-5}}$$

$$h = (10^{-5})^{\frac{1}{2}}$$

$$\log h = \frac{1}{2} \log 10^{-5}$$

$$= \frac{1}{2} \times -5 \log_{10} 10$$

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$\log h = -2.5$

$h = \text{Antilog}(-2.5)$

$= \text{Antilog}(1-0.5) \times 10^{-(2+1)}$

$= \text{Antilog}(0.5) \times 10^{-3}$

$h = 3.16 \times 10^{-3}$

$[H^+] = \sqrt{\frac{K_w \times K_a}{C}}$

$= \sqrt{\frac{10^{-14} \times 10^{-6}}{10^{-3}}} = \sqrt{10^{-17}} = 10^{-8.5}$

$pH = 8.5$

$pH = 7 + \frac{1}{2}pK_a + \frac{1}{2}\log C$

$= 7 + \frac{1}{2} \times 6 + \frac{1}{2} \times -3$

$= 10 - 1.5 = 8.5$

Q. Calculate K_h, h & pH of 10^{-4} M HCOOAg, given that K_a of formic acid 10^{-4} & K_b of $AgOH$ = 10^{-7}

→ HCOOAg,

WA WB,

$K_h = \frac{K_w}{K_a \times K_b} = \frac{10^{-14}}{10^{-4} \times 10^{-7}} = 10^{-3}$

$K_h = h^2$

$h = \sqrt{K_h} = \sqrt{10^{-3}} = 10^{-1.5}$

$\log h = -1.5$

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~~Antilog h~~ =

$h = \text{Antilog}(-1.5)$

$= \text{Antilog}(1-0.5) \times 10^{-(1+1)}$

$h = \text{Antilog}(0.5) \times 10^{-2}$

$\boxed{h = 3.16 \times 10^{-2}}$

$pH = 7 + \frac{1}{2}pK_a - \frac{1}{2}pK_b$

$= 7 + \frac{1}{2} \times 4 - \frac{1}{2} \times 7$

$= 9 - 3.5 = 5.5$

Q. $pK_h = pK_w - pK_a$ Eq^m is applicable for

(1) NaCl ✓(2) CH_3COONa (3) NH_4Cl (4) CH_3COONH_4.

→ $pK_w = pK_h + pK_a$

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

Q. Which of the following has max. value of h.

(1) NaCl (2) CH_3COONa (3) NH_4Cl ✓(4) CH_3COONH_4

Q. $CuSO_4$ aq. Sol^n is acidic in nature coz

✓(1) Presence of Cu^{+2} ions

(2) Presence of SO_4^{2-} ions

(3) Due to blue colour of sol^n

(4) Due to less solublity in water.

Q. If dissociation const. of a weaker base is $K_b = 10^{-5}$ ~~when it react with a strong base~~ calculate Eq^m constant when it react with strong acid.

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Neutralisation

$$\rightarrow \quad WB + SA \underset{\text{hydrolysis } K_h = ?}{\overset{K_N,\ K_f = ?}{\rightleftharpoons}} SAWB + H_2O$$

$$K_h = \frac{K_w}{K_b} = \frac{10^{-14}}{10^{-5}} = 10^{-9}$$

$$K_N = \frac{1}{K_h} = 10^{9}$$

Q: Which of the following has min. pH
(i) $NaCl$ (ii) CH_3COONa ✓(iii) NH_4Cl (iv) CH_3COONH_4

Q: Arrange the following solⁿ in increasing order of their pH
① 0.01 M H_2S (WA) ② 0.01 M HNO_3 (SA)
③ 0.01 M $BaSO_4$ (SASB) ④ 0.01 M Na_2S (WASB)
⑤ 0.01 M KOH (SB)

→ ② < ① < ③ < ④ < ⑤

· **Conjugate Acid-Base System:**

$$HCN \rightleftharpoons H^+ + CN^-$$

$$HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$$

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]} \quad \text{——— ①}$$

$$CN^- + H_2O \rightleftharpoons HCN + OH^-$$

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$$K_b = \frac{[HCN][OH^-]}{[CN^-]} \quad \text{—— (2)}$$

$$K_a \times K_b = [H_3O^+][OH^-]$$

$$\boxed{K_a \times K_b = K_w} = \overline{10^{-14}}$$

Q. Calculate diss. const. of ~~a~~ CH_3COOH if K_b of CH_3COO^- ions is 10^{-9}.

→

$$K_a \times K_b = K_w$$
$$K_a \times 10^{-9} = 10^{-14}$$
$$K_a = 10^{-5}$$

Q1 If K_a of NH_4^+ ion is 10^{-9}. Calculate K_b of NH_3.

Q2 If K_a of HCN 10^{-6}, then what would be K_b of Amonium Hydroxide.

→ ① 10^{-5} ② No relation b/w HCN & NH_4OH.
They're not conjugate acid-base.