Ionic Equilibrium

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

} Sorenson's equation

Abhishek

→ $H_2SO_4 \longrightarrow 2H^+ + SO_4^{--}$

- 2M, H_2SO_4 ⇒ 4N H_2SO_4

$[H^+]$ ⇒ Normality of acid

Q. 2M, H_2SO_4, $[H^+]$ = ?

N = 4N, So, $[H^+] = 4$

Acid + Acid

↓

→ $[H^+] = NR = \frac{N_1V_1 + N_2V_2 + \dots}{V_1 + V_2 + \dots}$

At 25°C,

$pH + pOH = 14$

$\therefore pH = 14 - pOH$

1. $\log 1 = 0$
2. $\log 2 = 0.3$
3. $\log 3 = 0.47$
4. $\log 4 = 0.6$
5. $\log 5 = 0.7$
6. $\log 6 = 0.77$
7. $\log 7 = 0.84$
8. $\log 8 = 0.9$
9. $\log 9 = 0.95$
10. $\log 10 = 1$
11. $\log 1.1 = 0.04$

* Acid / Acidic solution = $[H^+]$ = pH
* Base / Basic solution = $[OH^-]$ = pOH

$pOH = 14 - pH$

Q. $[H^+] = 2 \times 10^{-3}$ M

pH = ?

→ $pH = -\log[H^+]$

$= -\log 2 \times 10^{-3}$

$= 3 - \log 2$

$= 3 - 0.3$

$= 2.7$ Ans.

Q. $[H^+] = 5 \times 10^{-4}$ M, pH = ?

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

$= -\log 5 \times 10^{-4}$

$= 4 - \log 5$

$= 4 - 0.7$

$= 3.3$

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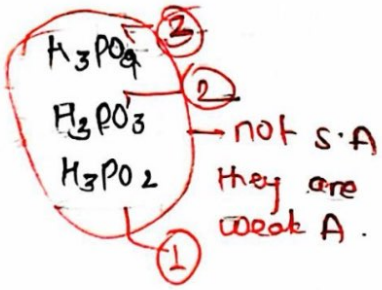
pH of strong acids

$HClO_4$, H_2SO_4, HNO_3, HCl, HBr, HI

1 2 1 1 1 1 → V.F



$[H^+]$ = Normality of acid = Molarity of acid × V.F

Q. pH = ?
10^{-4} M H_2SO_4. (molarity)

→ $[H^+]$ = normality
$= 2 \times 10^{-4} M$
$= -\log [H^+]$
$= -\log 2 \times 10^{-4}$
$= 4 - \log 2$
$= 4 - 0.3$
$[H^+] = 3.7$

Q. pH = ?
10^{-4} N H_2SO_4.

→ $[H^+]$ = normality = (4)
$= [10^{-4}]$
$= -\log 10^{-4}$
$= 4 - \log 1$
$= 4.$

pH of strong base

NaOH, KOH, $Ba(OH)_2$ -- alkali metals base, alkaline (except $Be(OH)_2$) Base.

$[OH^-]$ = Normality of Base = Molarity of BASE × V.F

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

$\therefore$ $pH = 14 - pH$

Q. 10^{-4} M $Ba(OH)_2$, pH = ?

$[OH^-] = 2 \times 10^{-4}$
$= -\log 2 \times 10^{-4}$
$= 4 - \log 2$
$= 4 - 0.3$
$pOH = 3.7$

$\therefore$ pH = 14 − p(OH)
= 14 − 3.7
= 10.3 Ans.

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... of $\frac{M}{500}$ $Ba(OH)_2$ soln:-

→ $N = \frac{M}{500} \times 2 = \frac{1}{250} = 4 \times 10^{-3}$

$pOH = -\log \frac{1}{250}$

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

$= 3 - \log 4$

$= 3 - 0.6$

$pOH = 2.4$

$\therefore pH = 14 - 2.4$

$\boxed{pH = 11.6}$

Q. In H_2SO_4, conc. of H^+ ion is 2×10^{-4} M, Ando out pH of solution:—

→ $[H^+] = 2 \times 10^{-4}$ M (not of H_2SO_4)

└ H^+ को conc. ही दिएको छ।

So, $pH = -\log 2 \times 10^{-4}$

$= 4 - \log 2$

$= 4 - 0.3$

$\boxed{pH = 3.7}$

Q. 0.0049 gm eq. H_2SO_4 is present in 10L solution. Then find out pH of soln:-

→ gm eq = 0.0049 , V = 10L

$N_{H_2SO_4} = \frac{\text{gm eq}}{V_L} = \frac{0.0049}{10} = 4.9 \times 10^{-4} N.$

$pH = -\log 4.9 \times 10^{-4}$

$= 4 - 0.7$

$\boxed{pH = 3.3}$ Ans:

Q. If NaOH is ~~that~~ in 500 mL solution. Find the pH of soln.

$\rightarrow$ $40 = m.w$

$$W = \frac{NEV}{1000}$$

$$1 = \frac{N \times \cancel{40} \times \cancel{500}}{\cancel{1000}}$$

$$N = \frac{1}{20} = 5 \times 10^{-2}$$

$$pOH = -\log 5 \times 10^{-2}$$
$$= 2 - \log 5$$
$$= 2 - 0.7$$
$$= 1.3$$

$$pH = 14 - 1.3$$
$$\boxed{pH = 12.7}$$

Q. In 250 mL solution, how many moles of $Ba(OH)_2$ should be that so that pH = 10.

so, $pOH = 14 - 10$
$= 4.$

$\because pOH = 4$

$-\log [OH^-] = 4$

N. of $Ba(OH)_2 = [OH^-] = 10^{-4}$

$$\therefore M = \frac{N}{2} = \frac{10^{-4}}{2} = 5 \times 10^{-5}$$

$$\because M = \frac{n}{V_{ml}} \times 1000$$

$$5 \times 10^{-5} = \frac{n}{\cancel{250}} \times \cancel{1000}\ 4$$

$$\boxed{n = 1.25 \times 10^{-5}}$$

Q. pH of 25cc ($\frac{N}{100}$ HCl) $\sim 10^{-2}$N

$$pH = -\log [H^+]$$
$$= -\log 10^{-2}$$
$$= 2.$$

Q. Find out pH of 2% w/v NaOH soln.

$\rightarrow$ $M = \frac{\% w/v \times 10}{mol\ wt} = \frac{2 \times 10}{40} = \frac{1}{2}$
$= 0.5$
$= 5 \times 10^{-1}$

$$pOH = -\log 5 \times 10^{-1}$$
$$= 1 - 0.7$$
$$= 0.3.$$

$$\boxed{pH = 14 - 0.3 = 13.7}$$

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Q. 10mL $\frac{M}{100}$ HCl + 20mL $\frac{M}{100}$ HNO_3 + 50 mL $\frac{M}{100}$ H_2SO_4 were mixed.

Then find out the pH of resulting soln :—

→ $N_R = \frac{10 \times \frac{1}{100} + 20 \times \frac{1}{100} + 50 \times \frac{1}{100} \times 2}{80}$

$N_R = [H^+] = \frac{1.3}{80} = 0.01$

$= 1 \times 10^{-2}$

$pH = -\log 1 \times 10^{-2}$

$2 - \log 1$

$= 2$

Q. 40mL of decinormal H_2SO_4 soln is mixed with 60 mL H_2O calculate pH.

→ Diluted soln = N = ?

$N_1V_1 = N_2V_2$

$\frac{1}{10} \times 40 = N_2 \times 100$

$N_2 = \frac{4}{100} = 4 \times 10^{-2} = [H^+]$

$pH = -\log 4 \times 10^{-2}$

$= 2 - \log 4$

$= 2 - 0.6 = 1.4.$

Q. pH of a solution is 3. It is diluted how many times so that pH becomes 6.

→ $\boxed{\text{New } [H^+] = \frac{\text{Old } [H^+]}{\text{times}}}$

$10^{-6} = \frac{10^{-3}}{x}$

$\therefore x = 1000$ times.

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Q. pH of a solution is 2. It is diluted 1000 times find out new pH.

→ New $[H^+]$ = $\frac{\text{old } [H^+]}{x} = \frac{10^{-2}}{1000} = 10^{-5}$

$\therefore$ **pH = 5**

Q. 4L of 10^{-4} M NaOH + 6L of 2×10^{-3} M $Ba(OH)_2$. pH = ?

→ $N_f = \frac{10^{-4} \times 4 + 2 \times 10^{-3} \times 2 \times 6}{10}$

$[OH^-] = 2.8 \times 10^{-3}$

$\therefore pOH = 3 - \log 2.8$

$= 3 - 0.47$

$= 2.53$

$pH = 14 - 2.53$

pH = 11.47.

Q. 1L of 10^{-4} M HCl + 2L of 2×10^{-4} M H_2SO_4 + 3L of 4L of 4×10^{-4} M $Ba(OH)_2$ + 3×10^{-4} M NaOH +

gm eq of acid = $\frac{1 \times 10^{-4} + 2 \times 2 \times 10^{-4} \times 2}{5}$

$= \frac{9 \times 10^{-4}}{5}$

gm eq of base = $\frac{4 \times 10^{-4} \times 2 + 3 \times 10^{-4} \times 3}{7}$

$= \frac{41 \times 10^{-4}}{7}$

$\therefore N_f = \frac{41 \times 10^{-4} - 9 \times 10^{-4}}{10} = 32 \times 10^{-5}$

$pH = \log 32 \times 10^{-5}$

$= 5 - \log 32$

$= 5 - \log 2^5$

$= 5 - 5 \times 0.3$

$= 5 - 1.5$

$\therefore pH = 14 - 3.5$

$\boxed{pH = 10.5}$

pH calculation of weak acid / weak base

Weak acid → HCN, CH_3COOH, HNO_2, ---etc

Ionization of monobasic weak acid

α = Degree of ionization

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

$t=0$	C	0	0
$t=eqm$	$C - C\alpha$	$C\alpha$	$C\alpha$

$$K_a = \frac{C\alpha \times C\alpha}{C - C\alpha} = \frac{C^2\alpha^2}{C(1-\alpha)} = \frac{C\alpha^2}{1-\alpha} \quad -(1)$$

(K_a ← ionization constant / dissociation const.)

$1 - \alpha \approx 1$ (α - very small)

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

$$\therefore \boxed{\alpha = \sqrt{\frac{K_a}{C}}} \quad -(2)$$

(C ← concentr.)

$$\boxed{\alpha \propto \frac{1}{\sqrt{C}} \propto \sqrt{V}}$$ ← dilution

on dilution, $C \downarrow$, $\alpha \uparrow$.

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

$$\therefore \boxed{[H^+] = C\alpha = \sqrt{K_a \times C}} \quad -(3)$$

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

$= -\log \sqrt{K_a \times C}$

$pH = -\frac{1}{2} pK_a - \frac{1}{2} \log C$ $\quad (pK_a = -\log K_a)$

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

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① $K_a = c\alpha^2$

② $\alpha = \sqrt{\frac{K_a}{c}}$

③ $[H^+] = c\alpha = \sqrt{K_a \times c}$ — c → normality of acid

④ $pH = \frac{1}{2}(pK_a - \log c)$

For weak base

① $K_b = c\alpha^2$

ⓑ $\alpha = \sqrt{\frac{K_b}{c}}$

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

④ $pOH = \frac{1}{2}(pK_b - \log c)$

$\therefore pH = 14 - pOH$

Q. $10^{-2}M$ CH_3COOH, $K_a = 4 \times 10^{-4}$

① $[H^+] = ?$

$\therefore [H^+] = c\alpha = \sqrt{K_a \times c}$ — (M = N)

$= \sqrt{4 \times 10^{-4} \times 10^{-2}}$

$= .2 \times 10^{-3}$

② $pH = ?$

$pH = 3 - \log 2$

$= 3 - 0.3 = 2.7.$

③ Deg. of ionization (α)

$\alpha = \sqrt{\frac{K_a}{c}} = \sqrt{\frac{4 \times 10^{-4}}{10^{-2}}} = 2 \times 10^{-2} = 0.2$ (DOI)

$= 0.2 \times 100\%$

$= 20\%$ (in %)

Q. $10^{-2}M$ NaOH, $\alpha = 0.2\%$, $pH = ?$

$\therefore \alpha = \frac{0.2}{100} = 2 \times 10^{-3}$ $\quad c = 10^{-2}M = 10^{-2}N.$

$\therefore K_a = \alpha^2 c$

$= 4 \times 10^{-6} \times 10^{-2} = 4 \times 10^{-5}$

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$[OH^-] = c\alpha$

$= 10^{-2} \times \frac{0.2}{100} = 2 \times 10^{-5}$

$pOH = 5 - \log 2$

$= 4.7$

$\therefore pH = 14 - 4.7$

$= 9.3$

Q. In 10L solution, how many gm of NH_4OH should be present so that pH of solution is 11. ($K_b = 10^{-5}$)

(a) 3.5g (b) 0.35g (c) 35g (d) 3.05g

→ $pH = 11$ So, $pOH = 3$

$K_b = 10^{-5}$

So, $[OH^-] = 10^{-3} = \sqrt{K_b \times C}$

$10^{-6} = K_b \times C$

$C = 10^{-1} = N = M$

$\therefore$ Conc → 10^{-1}

$M = \frac{w}{m.wt \times V}$

$0.1 = \frac{w}{35 \times 10}$

$\therefore w = 35g$

Q. How many gm CH_3COOH should be present in 10L solution.

$K_a = 10^{-5}$ & $pH = 4$.

→ $pH = 4$ →

So, $[H^+] = 10^{-4}$

$K_b = 10^{-5}$

$\sqrt{K_a \times C} = 10^{-4}$

$K_a \times C = 10^{-8}$

$\therefore C = 10^{-3} = M = N$

$\therefore M = \frac{w}{m.wt \times V}$

$10^{-3} = \frac{w}{60 \times 10}$

$\therefore w = 0.6$ gm

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4. 0.18 M $H_2C_2O_4$ solution, if $\alpha = 0.2\%$.

$[H^+] = c\alpha$

$= \underbrace{10^{-2}}_{N=2\times M} \times 2 \times \frac{0.2}{100}$

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

$pH = 5 - \log 4$

$= 5 - 0.6$

$= 4.4$

common ion effect

→ shown by weak electrolyte in presence of strong electrolyte having at least 1 ion common.

→ Weak electrolyte को ionization मा separation.

→ Degree of ionization (α) ↓es.

[All salts are strong electrolyte)

① NaCl, NaOH → both strong
X X

② HCN, CH_3COOH → both weak
X X

③ NH_4OH NH_4Cl → ~ shows.
weak bas. st. elect

④ HCN HCl → ~ shows
weak strong

⑤ HCN NaOH → but no common ion (XX).
weak strong

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

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

conc ↑es. so, rxn backwards.
dissociation
common ion effect

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Q. 100mL $\frac{M}{10}$ H_2SO_4 + 100mL $\frac{M}{20}$ KOH + 200mL $\frac{M}{10}$ NaOH, final volume is 1L. pH = ?

→ $N_R = \frac{100 \times \frac{1}{10} \times 2 + 100 \times \frac{1}{20} + \frac{1}{10} \times 200}{1000\,(ml)}$

$[OH^-] = \frac{25 - 20}{1000} = \frac{5}{1000} = 0.002\,N$

(base more)

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

$\therefore$ ~~$S = 0.002$~~

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

$= 3 - \log 5$

$= 3 - 0.7$

$= 2.3$

$pH = 11.7$

gm eq = Mole × V.F = N × V$_L$ = $\frac{W}{E}$

Q. 9.8g H_2SO_4, 100mL 1M HNO_3 and 0.1 Mole NaOH mixed & final volume made 1 Litre. pH = ?

→ gm eq. of H_2SO_4 = $\frac{W}{E}$

$= \frac{9.8}{49} = 0.2$

gm eq. of HNO_3 = $\frac{1 \times 100}{1000} = 0.1$ ($\because$ M = N)

gm eq. of NaOH = $0.1 \times 1 = 0.1$ (V.F.)

gm eq. of acid = 0.2 + 0.1 = 0.3.

gm eq of base = 0.1

$\therefore N_R = \frac{0.3 - 0.1}{1}$

$[H^+] = 0.2 = 2 \times 10^{-1}$

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

$= 1 - \log 2$

$= 1 - 0.3$

pH = 0.7 Ans

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Q. The following solution mixed:-

(1) 0.1 mole NaOH in 2L.

(2) 8g NaOH in 1L

(3) 0.3g eq. NaOH in 3l. pH = ?

→ gm eq of NaOH = $0.1 \times \frac{1}{1 \times f}$

= 0.1

gm eq of NaOH in 1L = $8/40 = 1/5 = 0.2$

gm eq. of NaOH in 3L = 0.3

$$N_R = \frac{0.1 + 0.2 + 0.3}{2 + 1 + 3} = \frac{0.6}{6} = 0.1$$

$= 10^{-1}$

$[OH^-] = 10^{-1}$

$\therefore pOH = 1 \quad \therefore pH = 14 - 1$

$= 13$ Ans

Q. In 100mL, 0.01N HCl solution, 900mL water is ~~mixed~~ added Then, find out pH of resulting soln:—

Sol. $M_1V_1 = N_2V_2$ → dilution का फार्मूला

$0.01 \times 100 = 1000 \times N$

$N = \frac{0.01}{10} = 1 \times 10^{-3} = [H^+]$

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

$= 3$ Ans

Q. pH of 100 litre solution is 5. How many litre of water should be evaporated so that pH becomes 2.

→ $\therefore$ pH of $H_2O = 5$

$\therefore [H^+] = 10^{-5}$

$\therefore N_1 = 10^{-5}$

$V_1 = 100\,L$.

after evap.

pH of $H_2O = 2$

$\therefore [H^+] = 10^{-2}$

$N_2 = 10^{-2}$

$\therefore 10^{-5} \times 100 = 10^{-2} \times V_2$

$V_2 = 10^{-3} \times 100$

$= 0.1\,l$

$\therefore$ water evaporated

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Q. pH of 10L solution is 3. Find out pH of 100mL of this solution:-

Sol:

(pH = 3, $[H^+] = 10^{-3}$) — pH will remain same of same solution.
└ So, pH is intensive

pH of very dilute strong acid

→ Answer of pH of this solution is less than 7 i.e. in between (6.7 to 7).

e.g. 10^{-8} M HCl, 10^{-7} M HNO_3, 10^{-9} M HCl

— water को conc नि बोलेंगे।

Q. pH = 10^{-8} M HCl solution. is?

Sol:

10^{-8} M HCl = 10^{-8} N HCl = $[H^+]$

Total $[H^+] = 10^{-8} + 10^{-7}$ (water नि add. गर्ने)
└ $10^{-8} \times 10$

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

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

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

$= 8 - \log 11$

$= 8 - 1.04$

$= 6.96$

∴ pH = 6.96

Q. pH of H_2SO_4 is 5. If is diluted 1000 times then pH of resulting solution is:-

→ (New H^+) = $\frac{\text{old } H^+}{\text{times}} = \frac{10^{-5}}{1000} = 10^{-8}$ N.
└ pH

So, total $[H^+] = 10^{-7} + 10^{-8}$

∴ pH = 6.96

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Q. Ionization constant (K_a) for CH_3COOH is 2×10^{-5}. Find out the degree of ionization of 0.1M CH_3COOH in presence of 0.1M HCL.

→

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

C 0 0

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

$H^+ = (C\alpha + 0.1)$ → less & neglect

$$HCL \longrightarrow H^+ + Cl^-$$

0.1 M 0.1M 0.1M

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} = \frac{C\alpha \times C\alpha}{C - C\alpha} \quad \frac{C\alpha \times 0.1}{(C - C\alpha)}$$

For weak electrolyte.

$$K_a = \frac{\alpha \times 0.1}{1 - \alpha} \quad (1 - \alpha \approx 1)$$

$\cancel{2 \times 10^{-5}}$ $\therefore K_a = \alpha \times 0.1$

$$\alpha = \frac{K_a}{0.1} = \frac{2 \times 10^{-5}}{0.1} = 2 \times 10^{-4} \text{ Ans}$$

Trick

$$\alpha = \frac{K_a}{[\text{conc. of strong elect}]}$$

Q. Find out degree of ionization of 0.02M NH_4OH in presence of 0.1 M $Ba(OH)_2$ solution. ($K_b = 10^{-5}$).

Soln.

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

C 0 0

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

$$\alpha = \frac{K_b}{[\text{conc. of strong ele}]}$$

$$Ba(OH)_2 \rightleftharpoons Ba^{+2} + 2OH^-$$

0.1 0.1 $0.1 \times 2 = 0.2$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]} = \frac{C\alpha \times 0.2}{C(1 - \alpha)} = \frac{\alpha \times 0.2}{1}$$

$(1 - \alpha) \Downarrow 1$

$$K_b = \alpha \times 0.2$$

$$\boxed{\alpha = 10^{-5}/.. = 5 \times 10^{-3} \text{ Ans}}$$

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Q. ~~Find out~~ In a solution, 1M CH_3COOH and 0.1M HCl are present. Find out the conc of acetate ion ($K_a = 10^{-5}$)

Sol.

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

C = 1 | 0 | 0

C − Cα | Cα | Cα

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

0.1 | 0.1 | 0.1

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} \quad \text{(acetate ion)}$$

$$\Rightarrow \frac{x \times 0.1}{C=(1)} = 10^{-5}$$

$$\Rightarrow x = \frac{10^{-5}}{0.1} \Rightarrow$$

$$x = 10^{-4}.$$

Q. In a solution, 0.08M CH_3COOH and 1M CH_3COONa are present. Then find out the conc of H+. ($K_a = 2 \times 10^{-5}$)

Sol.

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

C | 0 | 0

C − Cα | Cα | Cα

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

1 | 1 | 1

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

$$2 \times 10^{-5} = \frac{1 \times x}{C(1-\alpha)} \quad \text{(1)}$$

$$\Rightarrow 2 \times 10^{-5} = \frac{x}{0.08}$$

$$\therefore x = 1.6 \times 10^{-6}$$

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Q. In a solution, 0.1M HCN & 0.2M NaCN are present. Find out the total conc of ~~OH~~ ~~in solution~~ CN^- ion. ($\alpha = 0.02$)

Sol

$$HCN \rightleftharpoons H^+ + CN^-$$
$$C \quad 0 \quad 0$$
$$C - C\alpha \quad C\alpha \quad C\alpha$$

$$NaCN \longrightarrow Na^+ + CN^-$$
$$0.2 \quad 0.2 \quad 0.2$$

Total conc of CN^- = $C\alpha + 0.2$

= $0.1 \times 0.02 + 0.2$

= 0.202 Ans

Q. Ionisation constant (K_a) of $CH_3COOH = 2 \times 10^{-5}$. Find out α of 0.1M CH_3COOH in ~~presence~~ of 0.1M H_2SO_4.

Sol

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$
$$C \quad 0 \quad 0$$
$$C - C\alpha \quad C\alpha \quad C\alpha$$

$$H_2SO_4 \longrightarrow 2H^+ + SO_4^{--}$$
$$0.1 \quad 2 \times 0.1 \quad 0.1$$

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} = \frac{C\alpha \times 2 \times 0.1}{C(1-\alpha)}$$

$$K_a = \frac{\alpha \times 2 \times 0.1}{1}$$

$$\alpha = \frac{2}{2} \qquad \alpha = \frac{K_a}{0.2} = 10^{-4}$$

$$\alpha = \frac{K_a}{[C]_{\text{conc. of strong electrolyte}}}$$

$\alpha \propto \frac{1}{\sqrt{C_{conc.}}} \propto \sqrt{V_{dilute}}$

Q. How much water should be added into 1L of BOH (aq) solution to double its ionization?

→ $\alpha \propto \sqrt{V}$

$\frac{\alpha_2}{\alpha_1} = \sqrt{\frac{V_2}{V_1}}$

$\therefore 2 = \sqrt{\frac{V_2}{1}}$

$\therefore V_2 = 4L$.

water added = 4 − 1

= 3L

Q. Calculate pH of 0.1M CH_3COOH. $K_a = 2 \times 10^{-5}$.

Sol:

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

$= \sqrt{2 \times 10^{-5} \times 10^{-1}}$

$= \sqrt{2} \times 10^{-3}$

$pH = -\log \sqrt{2} \cdot 10^{-3}$

$= 3 - \frac{1}{2} \log 2$

$= 3 - \frac{1}{2} \times 0.3 = 2.85$

Q. To which of the following, Ostwald dilution law is applicable?

(a) NaOH (b) NH_4OH (weak) (c) LiOH (d) KOH

Q. The hydrogen ion concentration in weak acid of dissociation constant K_a & concn c is nearly equal to :-

(a) $\sqrt{\frac{K_a}{c}}$ (b) c/K_a (c) $K_a c$ (d) $\sqrt{K_a c}$ $[H^+] = c\alpha = \sqrt{K_a \times}$

Q. If α is degree of dissociation, the concn of a weak electrolyte & K_a the acid ionization constant then correct relation b/w α, c & K_a is:-

(a) $\alpha^2 = \sqrt{\frac{K_a}{c}}$ (b) $\alpha^2 = \sqrt{c/K_a}$ (c) $\alpha = \sqrt{\frac{K_a}{c}}$ (d) $\alpha = \sqrt{c/K_a}$

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Q. If 0.1M solution of HCN is 0.01% ionized. the ionization constant for HCN is :-

(a) 10^{-9} (b) 10^{-7} (c) 10^{-5} (d) 10^{-3}

$K_a = C\alpha^2$

$K_a = 0.1 \times \left(\frac{0.01}{100}\right)^2$

$K_a = 10^{-7}$

Q. The pH of 10^{-8} M HCl solution of HCl in water is :-

(a) 8 (b) -8 (c) between 7 & 8 (d) between 6 & 7

Q. 10^{-2} M CH_3COOH solution if its pka is 4.4. pH = ?

→

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

$= \frac{1}{2}(4.7 - \log 10^{-2})$

$= \frac{1}{2}(4.7 + 2)$

$= \frac{1}{2} \times 6.7 = 3.35$

Factors affecting degree of ionization

(1) Nature of solute :- strong → $\alpha = 1$

weak → $\alpha << 1$.

(2) Nature of solvent

(3) Conc^{r} of electrolyte. → strong → $\alpha = 1$ at any conc^{r}.

weak → $\alpha \uparrow$ on $\uparrow$ in dilution

$\alpha \downarrow$ on $\downarrow$ in dilution

0.001M HCl	=	0.01M HCl	=	1M HCl
α_1		α_2		α_3

$(\alpha_1 = \alpha_2 = \alpha_3)$ - strong elect.

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0.001M HCN α_1 | 0.01M HCN α_2 | 0.1M HCN α_3

$(\alpha_1 > \alpha_2 > \alpha_3)$ for weak electrolyte

Q. Ionization constant of acetate ion is 10^{-9}. Calculate K_a of CH_3COOH.

→ K_b of CH_3COO^- = 10^{-9} ; K_a of CH_3COOH = ?

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

↓ acid ← each other → C.B.

K_a K_b

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

$K_a = \frac{10^{-14}}{10^{-9}} = 10^{-5}$

$pK_a + pK_b = 14$

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

(base) K_b (c. acid) K_a.

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

acidic strength $\propto K_a \propto \frac{1}{pK_a}$

A	B
$pK_a = 5$	$pK_a = 3$

A > B (acidity)

Q. Find out pH of 10^{-3}M NH_4OH while ionization constant of $NH_4^+ = 10^{-9}$. (K_b)

Sol:

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

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

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

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

$= 10^{-4}$.

$\therefore pH = -\log 10^{-4}$

$= 4$.

$\therefore pH = 14 - 4$

$[OH^-] = 10$

Q. 3.5 g NH_4OH is present in 100 mL solution. Find out pH of solution. ($K_b = 10^{-5}$).

$$\underset{\text{a.b}}{NH_4OH} \rightleftharpoons \underset{\text{C.B}}{NH_4^+} + OH^-$$

$\because K_a \times K_b = 10^{-14}$

$K_a = 10^{-9}$

$\therefore K_a = 10^{-9}$

$\because$ [C] of $NH_4OH = \frac{3.5}{35} = \frac{0.1}{0.1\,L}$

$= 1$

$M = N = [OH^-] = \sqrt{K_b \times C}$

$= \sqrt{10^{-5} \times 1} = 10^{-2.5}$

$\therefore pOH = 2.5$

$\therefore pH = 14 - 2.5$

$\boxed{pH = 11.5}$

Solubility & solubility product

S
unit → mole/L

K_{sp}
unite → $(mol/L^{-1})^n$

$A_x B_y$

$\therefore n = (x+y)$

- $Al_2(SO_4)_3$ → $n = 2+3 = 5$
- $NaCl$ → $n = 2$.
- $CuSO_4$ → $n = 1+1 = 2$
- $AlPO_4$ → $n = 1+1 = 2$
- $CaCl_2$ → $n = 1+2 = 3$
- $Ca_3(PO_4)_2$ → $n = 3+2 = 5$
- $AgCl$ → $n = 2$. → unit of $K_{sp} = (mol L^{-1})^2$.

→ At a particular temp the maximum amount of a solute which can be dissolved in a fixed amount of solvent is called solubility of that solute.

→ Solubility is defined only at saturated point.

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On the solubility, solute (salt) can be divided into 3 categories:-

1. If $s > 0.1$ mol/L → soluble salt
2. If $0.1 > s > 0.01$ mol/L → partially soluble salt
3. If $s < 0.01$ mol/L → sparingle soluble salt → (SSS)

- In H_2O, NaCl is soluble but AgCl is SSS.

Eg. $A_x B_y \rightleftharpoons A_x B_{y(aq)} \xrightarrow{100\% \text{ ionize}} xA^{y+} + yB^{x-}$ ($A_x B_y$ (s), $A_x B_{y(aq)}$ (s), xA^{y+} (xs), yB^{x-} (ys) → solubility prod)

- $PbCl_2 \rightleftharpoons Pb^{2+} + 2Cl^-$
- $A_p B_{q(aq)} \longrightarrow pA^{q+} + qB^{p-}$
- $Mg_3N_2 \rightleftharpoons 3Mg^{2+} + 2N^{2-}$
- $MgSO_4 \rightleftharpoons Mg^{++} + SO_4^{--}$
- $AgCl \rightleftharpoons Ag^+ + Cl^-$

→ ionization

$$K_{sp} = [A^{y+}]^x [B^{x-}]^y$$

$$K_{sp} = (xs)^x (ys)^y$$

$$\boxed{K_{sp} = x^x y^y s^{x+y}}$$

$$K_3(PO_4) \rightleftharpoons 3K^+ + PO_4^{3-}$$

($K_3(PO_4)$: s; $3K^+$: 3s; PO_4^{3-}: s)

$$K_{sp} = [K^+]^3 [PO_4^{3-}]$$

$$= (3s)^3 (s)$$

$$= 27s^4$$

or, $K_{sp} = x^x y^y s^{x+y}$

$$= 3^3 1^1 s^4$$

$$= 27s^4.$$

$$\boxed{K_{sp} = [\text{mol } L^{-1}]^4}$$ unit.

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1) AB type of salt

e.g. NaCl, $CaCO_3$, $BaSO_4$, KCl, AgCl, $AlPO_4$,

$$AB_{(s)} \rightleftharpoons A^+ + B^-$$

solid(x) $\quad s \quad s$

$K_{sp} = [A^+][B^-]$

$= s \times s = s^2$

$\therefore \boxed{K_{sp} = s^2}$

unit $(mol/L)^2$

or, $K_{sp} = x^x y^y s^{x+y}$

$= 1^1 1^1 s^{1+1}$

$= s^2$

$\therefore \boxed{s = \sqrt{K_{sp}}}$

Common ion effect

Q. Solubility of AgCl is s mol/L.
Then find out solubility in presence of c mol/L $AlCl_3$ solution.

$$AgCl_{(s)} \rightleftharpoons Ag^+ + Cl^-$$

$s \quad s$

$K_{sp} = s^2$

$$AlCl_3 \rightleftharpoons Al^{+3} + 3Cl^-$$

$c \quad c \quad 3c$

let, solubility of AgCl due to common ion is s'

$K_{sp} = [Ag^+][Cl^-]$ $\quad Cl^- = s' + 3c$

c neglect.

$K_{sp} = s' \, 3c$

$\therefore \boxed{s' = \frac{K_{sp}}{3c}}$

Q. Solubility product for saturated solution of AgCl- is $8 \times 10^{-6} M^2$.
Then find out its solubility in presence of 0.01M NaCl.

Sol:-

$$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$$

$s' \quad s'$

$$NaCl \rightleftharpoons Na^+ + Cl^-$$

0.01 $\quad$ 0.01 $\quad$ 0.01

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$K_{sp} = [Ag^+][Cl^-]$ $\quad Cl^- = (s' + 0.01)$ (neg)

$= s' \times 0.01$

$8 \times 10^{-6} = s' \times 0.01$

$\therefore \boxed{s' = 8 \times 10^{-4}}$

– In presence of common ion, solubility decreases.

Q. The solubility product of AgCl is 4.0×10^{-10} at 298 K. The solubility of AgCl in 0.04M $CaCl_2$ is:-

Soln $AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$

s' $\quad s'$ – solubility in presence of $CaCl_2$.

$CaCl_2 \rightleftharpoons Ca^{++} + 2Cl^-$

0.04 $\quad$ 0.04 $\quad$ $2 \times 0.04 = 0.08$

$K_{sp} = [Ag^+][Cl^-]$ $\quad s' + 0.08$ (neg)

$K_{sp} = s' \times 0.08$

$s' = \frac{4 \times 10^{-10}}{0.08} = 5 \times 10^{-9}$.

Q. Solubility of saturated solution of AgCl is 4×10^{-6}. (S) Find out solubility in presence of 0.002M $BaCl_2$.

→ $AgCl \rightleftharpoons Ag^+ + Cl^-$

s' $\quad s'$

$BaCl_2 \rightleftharpoons Ba^{2+} + 2Cl^-$

0.002 $\quad$ 0.002 $\quad$ 2×0.002

$K_{sp} = [Ag^+][Cl^-]$

$16 \times 10^{-12} = s' \times 2 \times 0.002$

$s' = \frac{16 \times 10^{-12}}{0.004} = 4 \times 10^{-9}$ M

$K_{sp} = 4 \times 10^{-6}$

$s^2 = K_{sp}$

$s = 2 \times 10^{-3}$ (solubility in water)

$K_{sp} = s^2$

$= 16 \times 10^{-12}$

Q. S product of saturated solution of NaCl is $36 \times 10^{-12} M^2$. Find out its solubility in g/20L. (Imp)

→ $NaCl \rightleftharpoons Na^+ + Cl^-$

s s

$K_{sp} = [Na^+] \times [Cl^-]$

$36 \times 10^{-12} = s^2$

$s = 6 \times 10^{-6}$ mol/L

gm/L = $6 \times 10^{-6} \times 58.5$ (mw. of NaCl)

∴ g/20l = $6 \times 10^{-6} \times 58.5 \times 20$

= $12 \times 10^{-6} \times 585$ g/20L

Q. Solubility of saturated solution of AgCl is 1.435 mg/L. Find out the solubility product.

→ 1.435×10^{-3} g/l. = g/L.

AgCl → 143.5 (M.W) ∴ M = $\frac{1.435 \times 10^{-3}}{143.5} = 10^{-5}$ mol/L

∴ $s = 10^{-5}$ mol/L

$K_{sp} = s^2 = (10^{-5})^2 = 10^{-10} (mol/L)^2$

Q. S. product of AgCl = 4×10^{-10}. The solubility of AgCl in presence of 0.04M $CaCl_2$.

→ $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

(s) s' s'

$CaCl_2 \rightleftharpoons Ca^{++} + 2Cl^-(aq)$

0.04 0.04 2×0.04

$s' = \frac{K_{sp}}{0.04 \times 2}$

$= \frac{4 \times 10^{-10}}{0.04}$ g

for $CaCl_2$,

$K_{sp} = 1^1 2^2 s^{1+2}$

$= 4s^3$.

A_2B_3

$K_{sp} = 2^2 \times 3^3 \times 2^{2+3}$

$= 9108 s^5$

Q. Solubility product for saturated solution of $AgCl_{(s)}$ is $8 \times 10^{-6} M^2$. Then find out its solubility in presence of 0.01 M NaCl?

→

$$AgCl_{(s)} \rightleftharpoons Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$$
$$s' \qquad\qquad s'$$

$$NaCl \rightleftharpoons Na^{+} + Cl^{-}$$
$$0.01 \qquad 0.01 \qquad 0.01$$

$K_{sp} = [Ag^+][Cl^-]$

$8 \times 10^{-6} = s' \times 0.01$

$s' = \frac{8 \times 10^{-6}}{0.01} = 8 \times 10^{-4}$ M

(2) AB_2 / A_2B type of salt

e.g. $CaCl_2$, K_2SO_4, $FeCl_2$, $MgCl_2$,

$$AB_2 \rightleftharpoons A^{+2} + 2B^{-}$$
$$s \qquad\quad 2s$$

$K_{sp} = [A^{+2}][B^-]^2$

$= s \times (2s)^2$

$K_{sp} = 4s^3 (mol/L)^3$

or, $K_{sp} = x^x y^y s^{x+y}$

$= 1^1 2^2 s^{1+2}$

$= 4s^3$

Q. Solubility of saturated solution of $BaCl_2$ is $s = 4 \times 10^{-6}$ mol/L. Find out solubility of in presence of 0.01M $Ba(OH)_2$.

Solu:-

$$BaCl_2 \rightleftharpoons Ba^{++} + 2Cl^{-}$$
$$s \qquad\quad 2s$$

$K_{sp} = s(2s)^2$

$= 4s^3 = 4 \times 64 \times 10^{-18}$

$= 256 \times 10^{-18}$

$BaCl_2 \rightleftharpoons Ba^{+2} + 2Cl^-$ — in presence of common ion

$Ba(OH)_2 \rightleftharpoons Ba^{+2} + 2OH^-$

0.01 — 0.01 + s' — 0.01

$K_{sp} = [Ba^{+2}][Cl^-]^2$ — 0.01 + s' (neglect)

$2.56 \times 10^{-18} = 0.01 \times 4s'^2$

$s' = 8 \times 10^{-8}$ mol/L

Q. In saturated solution of Na_2CO_3, conc. of $Na^+ = 10^{-5}$. Calculate K_{sp}.

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

$2s$ (= 10^{-5}) — s

$\therefore 2s = 10^{-5}$

$s = 5 \times 10^{-6}$.

$\therefore K_{sp} = 4s^3$

$= 4 \times (5 \times 10^{-6})^3$

$= 500 \times 10^{-18}$

$K_{sp} = 5 \times 10^{-16}$

Q. S. product of saturated solution of $Mg(OH)_2$ is 4×10^{-12}. Calculate pH.

→ $Mg(OH)_2 \rightleftharpoons Mg^{++} + 2OH^-$

s — $2s$

$\therefore K_{sp} = 4s^3$

$4 \times 10^{-12} = 4s^3$

$s = 10^{-4}$ mol/Litre

$K_{sp} = 4s^3$

$s = 10^{-4}$ mol/litre

$[OH^-] = 2s = 2 \times 10^{-4}$.

$pH = 4 - \log 2$

$= 4 - 0.3$

$= 3.7$.

$\therefore pH = 14 - 3.7$

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③ AB_3 / A_3B type of salt

$K_{sp} = x^x y^y s^{x+y}$

$= 1^1 3^3 s^{1+3}$

$K_{sp} = 27s^4$

Q. $Ag_2(CrO_4)$

$Ag_2(CrO_4) \rightleftharpoons 2Ag^+ + CrO_4^{2-}$

$K_{sp} = [Ag^+]^2 [CrO_4^{2-}]^1$

Q. Mg_3N_2

$Mg_3N_2 \rightleftharpoons 3Mg^{2+} + 2N^{3-}$

$K_{sp} = [Mg^{2+}]^3 [N^{3-}]^2$

Q. $PbCl_2$

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

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

Q. $Al_2(SO_4)_3$

$Al_2(SO_4)_3 \rightleftharpoons 2Al^{+3} + 3SO_4^{-2}$

$K_{sp} = [Al^{+3}]^2 [SO_4^{-2}]^3$

Q. 0.0288 mg (SO_4^{--}) ions are present in 10mL saturated solution of $Al_2(SO_4)_3$. Find K_{sp} = ?

→

$SO_4^{--} = w = 0.0288$

$V = 10mL$.

$[SO_4^{--}] = M = \frac{0.0288 \times 10^{-3}}{(Mw) \times 96 \times 10} \times 1000$

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

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$Al_2(SO_4)_3 \rightleftharpoons 2Al^{+3} + 3SO_4^{-2}$

$2s \qquad 3s$

So, $\because 3s = 3\times10^{-5}$

$\therefore s = 10^{-5}$

$\therefore K_{sp} = (2s)^2 \times (3s)^3$

$= 108 \times s^5$

$= 108 \times (10^{-5})^5$

$\therefore \boxed{K_{sp} = 108\times10^{-25}}$

Q. Which salt is more soluble?

(a) AgCl ($K_{sp} = 4\times10^{-10}$), (b) $S = \sqrt{K_{sp}} = s^2$ $\therefore s = \sqrt{K_{sp}} = 2\times10^{-5}$ mol/lit

(b) $CaCl_2$ ($K_{sp} = 4\times10^{-12}$), $K_{sp} = 4s^3$ $s = \sqrt[3]{\frac{K_{sp}}{4}} = 10^{-4}$ mol/lit

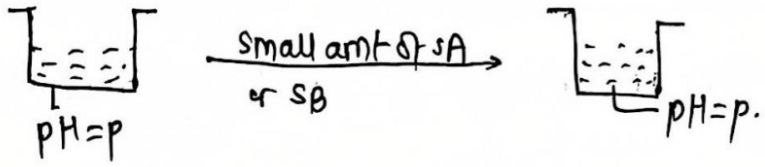
(c) $AlCl_3$ ($K_{sp} = 27\times10^{-24}$) $K_{sp} = 27s^4$ $s = 10^{-6}$ mol/lit.

(d) $PbCl_2$ ($K_{sp} = 4\times10^{-24}$) $K_{sp} = 4s^3$. $s = 10^{-8}$ mol/lit.

Solubility order → b > a > c > d

Buffer Solution

→ A solution that resist pH change on addition of small amount of strong acid or strong base.



* Types:-

① Simple buffer or Normal :-

- salt of weak acid and weak base

eg. MgF_2, CH_3COONH_4, NH_4CN.

$Mg(OH)_2 + HF \longrightarrow MgF_2$. $CH_3COOH + NH_4OH$ $NH_4OH + HCN$

(W.B) (W.A) W.A WB WB WA.

→ protein and amino acids are also simple buffer.

Nature:- Acidic, Basic, Neutral.

HF, HCl, HBr, HI
HF → W.A; HCl, HBr, HI → S.A

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

$Be(OH)_2$, $Mg(OH)_2$, $Ca(OH)_2$, ...
$Be(OH)_2$ → W.B; $Mg(OH)_2$, $Ca(OH)_2$, ... → S.B

② Mixed buffer

ⓐ Acidic buffer :-

– Weak ~~salt~~ acid and its salt with strong base.

e.g. ⓐ $HCN + NaCN$ ✓
HCN: W.A
NaCN → NaOH (S.B), HCN (W.A)

• $HCN + NaCl$ X
HCN: W.A; NaCl: salt of strong base & acid
X → not acidic buffer

✓ $CH_3COOH + CH_3COOK$ ✓
CH_3COOH: (W.A)
CH_3COOK → CH_3COOH, KOH (S.B)

$H_2S + NaHS$ ✓
H_2S: W.A.
NaHS → NaOH, H_2S.
NaOH: S.B

(2H dif.)
$H_2S + Na_2S$ X – but dif is not of single proton, 2 H are replaced.
Na_2S → NaOH, H_2S
So, not acidic buffer

$HClO + NaClO$ ✓
$HClO$: W.A
NaClO → NaOH, HClO
NaOH: S.B

$H_2CO_3 + NaHCO_3$ ✓

$NaHCO_3 + Na_2CO_3$ → Blood buffer

$N_3H + KN_3$
N_3H: W.A
KN_3 → KOH + HN_3
KOH: S.B

$HF + NaF$
HF: W.A
NaF → NaOH, HF
NaOH: S.B

$HCN + NaCN$
HCN: W.A
NaCN → NaOH, HCN
NaOH: S.B

$pH = pK_a + \log \frac{(salt)}{acid}$
↳ Henderson's eqn

→ Nature - Acidic.

→ pH < 7.

Blood buffer → $NaHCO_3 + Na_2CO_3$
↳ (acidic) - 7.3 pH

② Basic buffer

- weak base + its salt with strong acid.

eg. $NH_4OH + NH_4Cl$ ✓ , ✓ $Mg(OH)_2 + Mg(OH)Cl$

W.B, NH_4OH, HCl S.A, W.B, $Mg(OH)_2$, HCl S.A

$NaOH + NaCl$ — X

S.A

✓ $Mg(OH)Cl + MgCl_2$

W.B, $Mg(OH)_2$, HCl (S.A)

~~$pH = pKa + \log \frac{(salt)}{}$~~

→ Nature → Basic

→ pH > 7

$$pOH = pK_b + \log \frac{(salt)}{base}$$

$$pK_b = -\log K_b.$$

$$pH = 14 - pOH$$

Q. Find the pH of buffer solution containing 0.2 M CH_3COOH and 0.3 M CH_3COONa if Ka of CH_3COOH is 2×10^{-5}.

↳ Ka

→

$$pH = pKa + \log \frac{(salt)}{(acid)}$$

$$= 4.7 + \log \frac{0.3}{0.2}$$

$$= 4.7 + \log 3 - \log 2$$

$$= 4.7 + 0.47 - 0.3$$

$$= 4.87$$

$$pKa = -\log Ka$$

$$= 5 - \log 2$$

$$= 4.7.$$

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