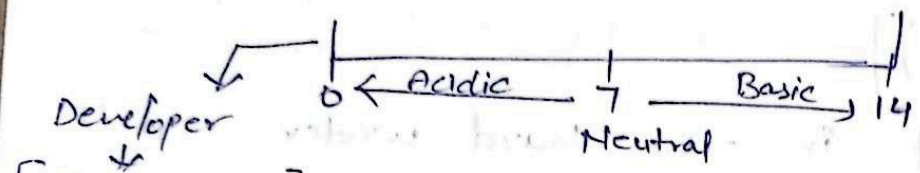
PH & Buffer solution

PH :- Potential / Power of Hydrogen.



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

PH of any solution can be defined as the -ve logarithm of its Hydrogen ion concentration.

$$[H^+][OH^-] = 10^{-14}$$

$$(10^{-7})(10^{-7}) = 10^{-14}$$

$$P_{OH} = -\log [OH^-]$$

$$[P_H = 14 - P_{OH}]$$

$$[H^+] = 10^{-PH}$$

.001 (H) HCl → $H = (H^+/OH^-)$

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

.001 .001 .001

.001M H_2SO_4

$H_2SO_4 \longrightarrow 2H^+ + SO_4^-$

$2 \times .001$

Q what will be the PH of following

i) .01 M HCl
ii) .02 N HNO_3
iii) .02 M H_2SO_4
iv) .0001 M NaOH
v) .02 M $Ba(OH)_2$
vi) .04 N H_2SO_4

(i) $[H^+] = .01 = \frac{1}{100} = 10^{-2} = -\log 10^{-2} = 2$

(ii) $[H^+] = .04 = 4 \times 10^{-2} = -\log(4 \times 10^{-2})$
$= -\log 4 + -\log 10^{-2}$

(iii) $H_2SO_4 \longrightarrow 2H^+$
$2 \times .02 = .04\ M\ (H^+)$

(iv) .0001 M NaOH
$[OH^-] = \frac{1}{10000} = 10^{-4}$

(v) $Ba(OH)_2 \longrightarrow 2OH^-$
$= 2 \times .02$
$= .04\ [OH^-]$

(vi) .04 N H_2SO_4
$N = M \times n\text{factor}$
$.04 = M \times 2$
$M = .02$

PH of 10^{-7} M HCl ⟶ In Acid PH is not equal or more than 7

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

$10^{-7} \quad 10^{-7}$

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

water $\hookrightarrow 10^{-7} \quad 10^{-7} \quad 10^{-7}$

So, $[H^+] = 10^{-7} + 10^{-7} = 2(10^{-7})$

$PH = -\log 2 + \log 7$

$PH = 6.69$

PH of 10^{-8} ~~NaOH~~ HCl

$[OH^+] = 10^{-8} + 10^{-7} = 10^{-7}\left(\frac{1}{10} + 1\right)$

$[OH^+] = 10^{-7} \Rightarrow 10^{-7}\left(\frac{11}{10}\right)$

$PH = -\log\left(10^{-7}\left(\frac{11}{10}\right)\right)$

$= 7 - \log(11 \times 10^{-7}) = 7 - \log 11 + 1$

~~$= 7 - \log 11 + \log 1$~~ $= 8 - \log 11$

$= 6.95$

~~$= 8.04$~~

Q what will be the pH of the Resultant sol^{n} of ~~50 ml of~~ after mixing it 50 ml of .2 M HCl mixing 50 ml of 0.1 M NaOH

$N_1V_1 \sim N_2V_2 = NV$
(Acid) (Base) (Resultant)

$.2 \times 50 \sim .1 \times 50 = N(50+50)$

$10 - 5 = 100N$

$N = .05$

$M = .05\ M$

$pH = 1.301$

Q calculate the pH of the resultant sol^{n} obtained by mixing ~~20~~25 ml of .2 M HCl and 50 ml of .25 M NaOH

$N_1V_1 \sim N_2V_2 = NV$

$.2 \times 25 \sim 50 \times .25 = N(25+50)$

$5 - 12.5 = N(75)$

$N = 0.1$

$M = \frac{N}{n\text{factor}} = 0.1\ M$

$pOH = -\log[OH^-] = -\log[0.1]$

$pOH = 1$

$pH = 14 - 1 = 13$

Buffer solution

A solution which mantain its pH value fairly constant even on the addition of small amount of acid and base such sol^{n} are called Buffer sol^{n}.

(OR)

A sol^{n} which resist any change with its pH is called as Buffer solution.

Acidic Buffer:- Mixture of weak Acid and its salt with a strong Base

eg:- $CH_3COOH + CH_3COONa$.

Basic Buffer:- Mixture of weak Base and its salt with a strong Acid.

eg:- $NH_4OH + NH_4Cl$

Hendersen Equation for finding out the pH of Acidic Buffer solution.

V.V.V. V.Imp

$$CH_3COOH \underset{}{\overset{Ka}{\rightleftharpoons}} CH_3COO^- + H^+$$

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

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

On Adding Sodium Acetate the degree of ionisation of Acetic Acid gets supressed due to a common Acetate ion in both as such we assume that whole of the acetate ion in the solution will come from the ionisation of Sodium Acetate.

Therefore, the eq① can be written as :-

$$K_a = \frac{[CH_3COONa][H^+]}{[CH_3COOH]}$$

$$\left[K_a = \frac{[salt]}{[Acid]} [H^+] \right]$$

(or)

$$\left[[H^+] = K_a \times \frac{[Acid]}{[salt]} \right]$$

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

$$\left\{ P_H = -\log K_a + \log \left[\frac{salt}{Acid} \right] \right\}$$

$$\left[P_H = P_{Ka} + \log \frac{[salt]}{[Acid]} \right]$$

I)
* If some external acid is added to the acidic buffer :-

$$P_H = -\log K_a + \log \frac{[salt] - [acid_{ext.}]}{[Acid] + [acid_{ext.}]}$$

Q Imp Find the pH of buffer sol^n containing .2 mole/L of sodium Acetate and .5 moles/L of CH_3COOH. The value of Ka for is 1.8×10^{-5}

$$P_H = -\log Ka + \log \frac{[salt]}{[Acid]}$$

$$P_H = -\log (1.8 \times 10^{-5}) + \log \frac{[.2]}{[.5]}$$

$$[P_H = 4.346]$$

II) If Ext. base is added to the acidic Buffer sol^n:

$$\left[P_H = -\log Ka + \log \frac{[salt] + [base_{ext}]}{[Acid] - [base_{ext}]} \right]$$

Q a) A litre of a sol^n contains .1 mole of Acetic Acid and .1 mole of sodium Acetate calculate the change in pH of buffer .02 mole of NaOH added to this sol^n. $Ka = 1.8 \times 10^{-5}$

$$P_{H_1} = P_{ka} + \log \frac{[.1]}{[.1]}$$

$$P_{H_2} = P_{ka} + \log \frac{[.1] + [.02]}{[.1] - [.02]}$$

$$\Delta PH = P_{H_2} - P_{H_1}$$

b) Add 100 ml of .1 mole HCl → neglect external Acid / Base volume