HYDROGENION CONCENTRATION-pH SCALE

It is clear from the above discussion that nature of the solution (acidic, alkaline or neutral) can be represented in terms of either hydrogen ion concentration or hydroxyl ion concentration but it is convenient to express acidity or alkalinity of a solution by referring to the concentration of hydrogen ions only. Since H^+ ion concentration can vary within a wide range from 1 mol per litre to about 1.0×10^{-14}mol per litre, a logarithmic notation has been devised **by Sorensen**, in 1909, to simplify the expression of these quantities. The notation used is termed as the pH scale.

The hydrogen ion concentrations are expressed in terms of the numerical value of negative power to which 10 must be raised. This numerical value of negative power was termed as pH, *i.e.*,

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

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

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

or $pH = \log 1/[H^+]$

pH of a solution is, thus, defined as **the negative logarithm of the concentration (in mol per litre) of hydrogen ions which it contains or pH of the solution is the logarithm of the reciprocal of H^+ ion concentration.**

Just as pH indicates the hydrogen ion concentration, the pOH represents the hydroxyl ion concentration, i.e.,

pOH = -log [OH-] Considering the relationship,

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

Taking log on both sides, we have

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

or $-\log [H^+] - \log [OH^-] = -\log K_w = -\log (1 \times 10^{-14})$

or $pH + pOH = PK_w^* = 14$

i.e., sum of pH and pOH is equal to 14 in any aqueous solution at 25°C. The above discussion can be summarised in the following manner:

	$[H^+]$	$[OH^-]$	pH	pOH
Acidic solution	$>10^{-7}$	$<10^{-7}$	<7	>7
Neutral solution	10^{-7}	10^{-7}	7	7
Basic solution	$<10^{-7}$	$>10^{-7}$	>7	<7

$[H^+]$	$[OH^-]$	pH	pOH	Nature of solution
10^{0}	10^{-14}	0	14	Strongly acidic Acidic
10^{-2}	10^{-12}	2	12	Weakly acidic
10^{-5}	10^{-9}	5	9	Neutral
10^{-7}	10^{-7}	7	7	Weakly basic
10^{-9}	10^{-5}	9	5	Basic
10^{-11}	10^{-3}	11	3	Strongly basic
10^{-14}	10^{-0}	14	0	

The following table shows the pH range for a few common substances:

Substance	pH range	Substance	pH range
Gastric contents	1.0-3.0	Milk (cow)	6.3-6.6
Soft drinks	2.0-4.0	Saliva (human)	6.5-7.5
Lemons	2.2-2.4	Blood plasma (human)	7.3-7.5
Vinegar	2.4-3.4	Milk of magnesia	10.5
Apples	2.9-3.3	Sea water	8.5
Urine(human)	4.8-8.4		

Any method which can measure the concentration of H^+ ions or OH^- ions in a solution can serve for finding pH value.

Limitations of pH Scale:

(i) pH values of the solutions do not give us immediate idea of the relative strengths of the solutions. A solution of pH = 1 has a hydrogen ion concentration 100 times that of a solution of pH = 3 (not three times). A 4×10^{-5} ***N*** HCI is twice concentrated of a 2×10^{-5} ***N*** HCI solution, but the pH values of these solutions are 4.40 and 4.70 (not double).

(ii) pH value of zero is obtained in 1 A' solution of strong acid. In case the concentration is 2 ***N***, 3 ***N***, 10 ***N***, etc. The respective pH values will be negative.

(iii) A solution of an acid having very low concentration, say 10^{-8} ***N***, cannot have pH 8, as shown by pH formula, but the actual pH value will be less than 7.

Note:

(i) Normality of strong acid = $[H_3O^+]$

Normality of strong base = $[OH^-]$

$\therefore$ pH = -log [N] for strong acids

pOH = -log [N] for strong acids

(ii) Sometimes pH of acid comes more than 7 and that of base comes less than 7. It shows that the solution is very dilute; in such cases, H^+ or OH^- contribution from water is also considered, e.g., in 10 ***N*** HCI,

$[H^+]_{Total} = [10^{-8}]_{Acid} + [10^{-7}]_{Water}$

$= 11 \times 10^{-8}$ M$= 1.1 \times 10^{-7}$ M

(iii) **pH of mixture:**

Let one litre of an acidic solution of pH 2 be mixed with two litre of other acidic solution of pH **3**. The resultant pH of the mixture can be evaluated in the following way.

Sample 1	**Sample 2**
pH = 2	pH = 3
$[H^+]=10^{-2}$ M	$[H^+] = 10^{-3}$M
V = 1 litre	V= 2 litre

$M_lV_l+M_2V_2 = M_R(V_i + V_2)$ $10^{-2} \times 1 + 10^{-3} \times 2 = M_R$ (1 +2)

$(12 + 10^{-3})/3 = M_R$

$4 \times 10^{-3} = M_R$(Here, M_R = **Resultant molarity**) pH = -log (4 × 10^{-3})

(iv) Total concentration of $[H^+]$ or **$[H_3{}^+O]$** in a mixtureof **weak acid** and a **strong acid**.

$= (C_2 + \sqrt{(c^2{}_2 + 4K_aC_1)})/2$

where **C_1** is the concentration of weak acid (in mol litre having dissociation constant**K_a**

C_2 is the concentration of strong acid

(v) Total $[OH^-]$ concentration in a mixture of **two weak bases.**

$= \sqrt{(K_1C_1 + K_2C_2)}$

where ***K_1*** and ***K_2*** are dissociation constants of two weak bases having C_1 and C_2 as their mol $litre^{-1}$ concentration respectively.

Source : http://ciseche10.files.wordpress.com/2013/12/ionic-equilibrium.pdf