pH – A Measure of Hydrogen Ion Activity

We need to measure the concentrations of H^+ and OH^- ions in aqueous solutions. The problem here is that the concentrations are usually very small (range from $1.0x10^{-1}$ – $1.0x10^{-14}$) and hence it becomes inconvenient to work with such small negative exponent numbers.

To circumvent this difficulty, Danish chemist Soren Sorensen in 1909 proposed a more practical method of measuring the H^+ ion concentration and called it a pH. The pH of a solution is defined as *the negative logarithm of hydrogen ion concentration (in mol/L):*

$$pH = -\log[H^+] \qquad or \qquad pH = -\log[H_3O^+]$$

This is nothing but the transformation from one domain (concentration) to another (pH) domain. When this transformation applied, the small negative exponents of hydrogen ions concentrations become positive manageable pH numbers. In addition, the pH , like equilibrium constant, is also a dimensionless quantity, i.e., no units.

Acidic and basic solutions are distinguished using either H^+ concentrations or pH. Since the pH offers a simple way of expressing the H^+ ion concentration, acidic and basic solutions at 25^0C are distinguished by their pH values. Thus

Acidic solution: **pH < 7,** **$[H^+] > 1.0 \text{ x } 10^{-7}$ M**

Neutral solution: **pH = 7,** **$[H^+] = 1.0 \text{ x } 10^{-7}$ M**

Basic solution: **pH > 7,** **$[H^+] < 1.0 \text{ x } 10^{-7}$ M**

Sorensen's concept can also be applied to OH^- ion concentration and define **pOH** analogous to pH. It is defined as *the negative logarithm of hydroxide ion concnentration (mol/L).* Thus

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

Now let us apply this concept to ion-product (K_w) of water at 25^0C and express K_w in terms pf pH and pOH:

$$K_w = [H^+][OH^-] = 1.0x10^{-14}$$

Taking the negative logarithm of both sides, we obtain

$$-\log[H^+] + (-\log[OH^-]) = -\log(1.0x10^{-14})$$

or

$$-\log[H^+] + (-\log[OH^-]) = 14.0$$

Substituting the definition of pH and pOH, we derive the following equation:

$$pH + pOH = 14.0$$

This equation provides another relationship between the hydrogen ion concentration and hydroxide ion concentration.

The following is the pH scale that explains acid and base ranges and associated hydrogen and hydroxide concentrations.

pH Scale

pH	**$[H^+]$**	**pOH**	**$[OH^-]$**	**Range**
0	$1x10^{0}$	14	$1x10^{-14}$	
1	$1x10^{-1}$	13	$1x10^{-13}$	
.	.	.	.	acid range
.	.	.	.	
.	.	.	.	
7	$1x10^{-7}$	7	$1x10^{-7}$	neutral
.	.	.	.	
.	.	.	.	
.	.	.	.	base range
12	$1x10^{-12}$	2	$1x10^{-2}$	
13	$1x10^{-13}$	1	$1x10^{-1}$	
14	$1x10^{-14}$	0	$1x10^{0}$	

Remember :

- Lower the pH means higher the acidity and lower the pOH means higher the basicity. This tendency shown by arrows.
- pH + pOH =14
- K_w=$[H^+]$ $[OH^-]$ = $(1x\ 10^{-7})$ $(\ 1\ x10^{-7})$ = $1\ x10^{-14}$

You can verify these statements using the above pH scale.

So far, we have discussed how to get the pH or pOH knowing the hydrogen ion or hydroxide concentration. Suppose, we want get hydrogen ion or hydroxide concentration knowing pH or pOH, how would we do it? It is a simple matter that can be accomplished through the following equations.

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

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

The following table lists pH of some common substance that we come across.

Substance	**pH**
1.00 M HCl	0.0
Gastric juice	1.0 -2.0
Lemon juice	2.4
Vinegar	3.0
Grapefruit juice	3.2
Orange juice	3.5
Wine	3.5
Tomato juice	4.1
Urine	4.8-7.7
Black coffee	5.0
Saliva	6.4-6.9
Milk	6.5
Pure water	7.0
Blood	7.35-7.45
Tears	7.4
Sea water	8.5
Household detergent	9.2
Milk of magnesia	10.6
Household ammonia	115.5
1.00 M NaOH	14.0

Can you tell which is acid, which is base, and which is neutral based on the pH values?

Example - Calculating pH from H^+ ion concentration

Calculate the pH of

(a) 0.005 M H^+
(b) 3.5 M H^+

Answer

(a) $[H^+]$ =0.005M
pH = -log(0.005) = - (-2.3) = 2.3

(b) $[H^+]$ =3.5 M
pH = -log(3.5) = - 0.54

Example- Calculating pH from OH^- ion concentration

Calculate the pH of 0.15 M OH^-

Answer

This problem can be solved in two different ways:

i. $[OH^-] \rightarrow [H^+] \rightarrow$ pH

$[OH^-] = 0.15M$

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0x10^{-14}}{0.15} = 6.67x10^{-14}$$

pH = $-\log(6.67x10^{-14}) = 13.12$

ii. $[OH^-] \rightarrow$ pOH $\rightarrow$ pH

$[OH^-] = 0.15M$

pOH = -log(0.15M) = -(-0.82) =0.824

pH = 14.0 – pOH = 14.0 -0.824 = 13.12

Example – Calculating the concentration of hydrogen ion from pH

Calculate the hydrogen ion concentration and hydroxide ion concentration in milk, the pH of which is 6.5.

Answer

pH = 6.5

$[H^+] = 10^{-pH} = 10^{-6.5} = 3.2 \text{ x } 10^{-7}$ M

The OH^- ion concentration can be calculated in two different ways:

i. pH $\rightarrow$ pOH $\rightarrow [OH^-]$

pOH = 14.0 –pH = 14.0 -6.5 = 7.5

$[OH^-] = 10^{-pOH} = 10^{-7.5} = 3.2 \text{ x } 10^{-8}$ M

ii. pH $\rightarrow [H^+] \rightarrow [OH^-]$

$[H^+] = 10^{-pH} = 10^{-6.5} = 3.2 \text{ x } 10^{-7}$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0x10^{-14}}{3.2x10^{-7}} = 3.2x10^{-8}$$