THE pH SCALE

Acidic strength refers to an acid's inclination to yield H_3O^+ or H^+ ions when dissolved in water. Therefore, the greater the propensity to release H^+ ions, the stronger the acidity of the substance. On the other hand, basic strength pertains to a base's ability to provide OH^- ions when mixed with water.

So, greater the tendency to give OH^- ions, more will be basic strength of the substance. The concentration of H^+ ions is written in a simplified form introduced by Sorenson known as pH scale.

pH is defined as negative logarithm of activity of H^+ ions. activity (a) $= \gamma \cdot \frac{C}{C_0}$

C → unitless, C_0 → standard concentration

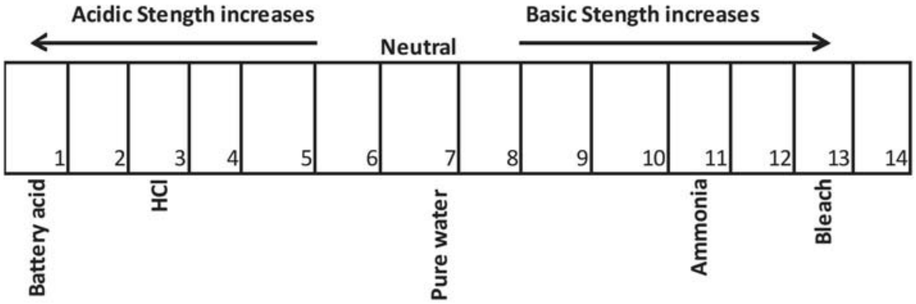
$$a = \gamma \cdot C$$

for dilute solution
$\gamma = 1$

for concentrated solution
$\gamma \neq 1$

$\therefore$ $a = C = [H^+]$

$\therefore$ pH = – log (where is the activity of H^+ ions)



The activity of H^+ ions represent the concentration of unbound H^+ or H_3O^+ ions within a solution. The pH scale, established from 0 to 14, designates a midpoint at 7 when using water as the solvent at 25 °C. It's essential to recognize that alterations in temperature and solvent will result in a modification of the pH scale's range.

For example,

0 - 14 at 25 °C ($K_W = 10^{-14}$) Neutral point, pH = 7

0 - 13 at 80 °C ($K_W = 10^{-13}$) Neutral point, pH = 6.5

pH can also be negative or > 14

Relation Between pH and pOH

The pH (potential of hydrogen) is determined by the negative logarithm (base 10) of the concentration of hydronium ions (H_3O^+), expressed as $pH = -\log[H_3O^+]$. Similarly, the pOH (potential of hydroxide) is calculated as the negative logarithm (base 10) of the concentration of hydroxide ions (OH^-), denoted as $pOH = -\log[OH^-]$.

The product of the concentrations of hydronium and hydroxide ions in a solution, represented as $[H_3O^+][OH^-]$, is equal to the ion product constant (Kw), which is 10^{-14} at 25°C. Mathematically, this is expressed as $[H_3O^+][OH^-]=K_w=10^{-14}$.

Taking the logarithm of both sides of the equation, we get $[H_3O^+] + \log[OH^-] = \log K_w = \log 10^{-14} = -14$. This equation signifies the relationship between the concentrations of hydronium and hydroxide ions and their connection to the ion product constant.

Furthermore, the sum of the pH and pOH in a solution is equal to the negative logarithm (base 10) of the ion product constant, denoted as pKw, which is 14 at 298 K. In other words, $pH + pOH = pKw = 14$ represents the equilibrium between the acidic and basic components in a solution at a given temperature.

Common Ion Effect

In the case of pure water, the concentrations of hydronium (H_3O^+) and hydroxide (OH^-) ions are both equal to 10^{-7} M. When an acid is introduced to pure water, the concentration of hydronium ions ($[H_3O^+]$) becomes greater than 10^{-7}, disrupting the equilibrium.

Conversely, when a base is added to pure water, the concentration of hydroxide ions ($[OH^-]$) exceeds 10^{-7}, while the concentration of hydronium ions ($[H_3O^+]$) is determined by Kw divided by the concentration of hydroxide ions, represented as $[H_3O^+] = K_w/[OH^-]$.

The concept of the common ion effect comes into play when a weak electrolyte's ionization is hindered in the presence of another electrolyte, particularly a strong electrolyte that shares a common ion with the weak electrolyte. This effect occurs because the addition of a common ion reduces the extent to which the weak electrolyte dissociates, impacting its ionization process.