# IONIC PRODUCT OF WATER

Pure water is a very weak electrolyte and ionises according to the equation

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

Applying law of mass action at equilibrium, the value of dissociation constant, K comes to

$K = [H^+] [OH^-]/[H_2O]$

or $[H^+][OH^-] = K[H_20]$

Since dissociation takes place to a very small extent, the concentration of undissociated water molecules, $[H_20]$, may be regarded as constant. Thus, the product #$[H_20]$ gives another constant which is designated as $K_w$. So,

$[H^+][OH^-] = K_w$

The constant, $K_w$, is termed as ionic product of water.

The product of concentrations of H1 and OH ions in water at a particular temperature is known as ionic product of water. The value of $K_w$ increases with the increase of temperature, i.e., the concentration of $H^+$ and $OH^-$ ions increases with increase in temperature.

| Temperature (°C) | Value of $K_w$ |
|---|---|
| 0 | $0.11 \times 10^{-14}$ |
| 10 | $0.31 \times 10^{-14}$ |
| 25 | $1.00 \times 10^{-14}$ |
| 100 | $7.50 \times 10^{-14}$ |

The value of $K_w$ at 25°C is $1 \times 10^{-14}$. Since pure water is neutral in nature, $H^+$ ion concentration must be equal to $OH^-$ ion concentration.

$[H^+] = [OH^{\sim}] = x$

or $[H^+][OH^-]=x^2= 1 \times 10^{-14}$

or $x = 1 \times 10^{-7}$ *M*

or *$[H^+] = [OH^-] = 1 \times 10^{-7}$ mol litre$^{-1}$*

This shows that at 25°C, in 1 litre only $10^{-7}$ mole of water is in ionic form out of a total of approximately 55.5 moles.

When an acid or a base is added to water, the ionic concentration product, $[H^+][OH^-]$, remains constant, i.e., equal to $K_w$ but concentrations of $H^+$ and $OH^-$ ions do not remain equal. The addition of acid increases the hydrogen ion concentration while that of hydroxyl ion concentration decreases, i.e.,

$[H^+] > [OH^-]$; (Acidic solution)

Similarly, when a base is added, the $OH^-$ ion concentration increases while $H^+$ ion concentration decreases,

i.e., $[OH^-] > [H^+]$; (Alkaline or basic solution)

In neutral solution, $[H^+] = [OH^-] = 1 \times 10^{-7}$ *M*

In acidic solution, $[H^+] > [OH^-]$

or $[H^+] > 1 \times 10^{-7}$ *M*

and $[OH^-] < 1 \times 10^{-7}$ *M*

In alkaline solution, $[OH^-] > [H^+]$

or $[OH^-] > 1 \times 10^{-7}$ *M*

and $[H^+] < 1 \times 10^{-7}$ *M*

Thus, if the hydrogen ion concentration is more than $1 \times 10^{-7}$ M, the solution will be acidic in nature and if less than $1 \times 10^{-7}$ M, the solution will be alkaline.

$[H^+] = 10^{-0}\ 10^{-1}\ 10^{-2}\ 10^{-3}\ 10^{-4}\ 10^{-5}\ 10^{\sim 6}$ (Acidic)

$[H^+] = 10^{-7}$ (Neutral)

$[H^+] = 10^{-14}\ 10^{-13}\ 10^{-12}\ 10^{-11}\ 10^{-10}\ 10^{-9}\ 10^{-8}$ (Alkaline)

We shall have the following table if $OH^-$ ion concentration is taken into account.

$[OH^-] = 10^{-14}\ 10^{-13}\ 10^{-12}\ 10^{-11}\ 10^{-10}\ 10^{-9}\ 10^{-8}$ (Acidic)

$[OH^-] = 10^{-7}$ (Neutral)

$[OH^-] = 10^{-0}\ 10^{-1}\ 10^{-2}\ 10^{-3}\ 10^{-4}\ 10^{-5}\ 10^{-6}$ (Alkaline)

It is, thus, concluded that every aqueous solution, whether acidic, neutral or alkaline contains both $H^+$ and $OH^-$ ions. The product of their concentrations is always constant, equal to $1 \times 10^{-14}$ at 25°C. If one increases, the other decrease accordingly so that the product remains 1×10-14 at 25° C.

If $[H^+] = 10^{-2}$ M, then $[OH^-] = 10^{-12}$ M; the product, $[H^+][OH^-] = 10^{-2} \times 10^{-12} = 10^{-14}$; the solution is acidic.

**If $[H^+] = 10^{-10}$ M, then $[OH^-] = 10^{-4}$ M; the product, $[H^+][OH^-] = 10^{-10} \times 10^{-4} = 10^{-14}$; the solution is alkaline.**

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