Ionization of Water

$$H_2O(l) + H_2O(l) \leftrightarrows H_3O^+(aq) + OH^-(aq)$$

The self-ionization of water occurs to a very small extent. All water samples will contain both the hydronium ion and the hydroxide ions. In pure water the concentration of each of these ions is equal (1×10^{-7} M), therefore pure water is neutral. An acidic or basic solution exists when the concentration of one is greater than the other.

$$K_w = [H_3O^+][OH^-] = 1 \times 10^{-14}$$

K_w - the ionization constant for water

In any aqueous solution the concentrations of hydroxide and hydronium ions are interdependent. As the concentration of one increases the other decreases.

The pH scale - 1909

0 ----------------------------------- 7 ---------------------------------- 14

very acidic	slightly acidic	↓ ↓ neutral	slightly basic	very basic

Calculation of pH

$pH = -\log[H_3O^+]$ $\qquad$ $pOH = -\log[OH^-]$

$[H_3O^+] = \text{antilog}(-pH)$ $\qquad$ $[OH^-] = \text{antilog}(-pOH)$

$[H_3O^+] = 10^{-pH}$ $\qquad$ $[OH^-] = 10^{-pOH}$

$$pH + pOH = 14$$

For each of the following, calculate $[H_3O^+]$, $[OH^-]$, pH, and pOH.

1. $[H_3O^+] = 1.56 \times 10^{-3}$ mol/L
2. $[H_3O^+] = 0.125$ M
3. $[OH^-] = 4.12 \times 10^{-4}$ mol/L
4. $[OH^-] = 0.0250$ M
5. pH = 3.45
6. pH = 9.76
7. pOH = 2.45
8. pOH = 12.5