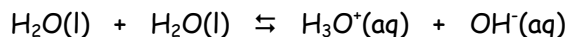


Ionization of Water



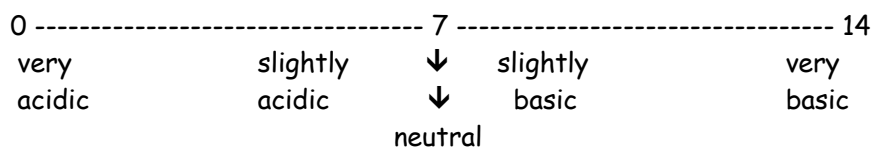
The self-ionization of water occurs to a very small extent. All water samples will contain both the hydronium ion and the hydroxide ions. In pure water the concentration of each of these ions is equal ($1 \times 10^{-7} \text{ M}$), therefore pure water is neutral. An acidic or basic solution exists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

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$$

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



Calculation of pH

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \text{antilog}(-\text{pH})$$

$$[\text{OH}^-] = \text{antilog}(-\text{pOH})$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

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

$$\text{pH} + \text{pOH} = 14$$

For each of the following, calculate $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, pH, and pOH.

1. $[\text{H}_3\text{O}^+] = 1.56 \times 10^{-3} \text{ mol/L}$
2. $[\text{H}_3\text{O}^+] = 0.125 \text{ M}$
3. $[\text{OH}^-] = 4.12 \times 10^{-4} \text{ mol/L}$
4. $[\text{OH}^-] = 0.0250 \text{ M}$
5. $\text{pH} = 3.45$
6. $\text{pH} = 9.76$
7. $\text{pOH} = 2.45$
8. $\text{pOH} = 12.5$