

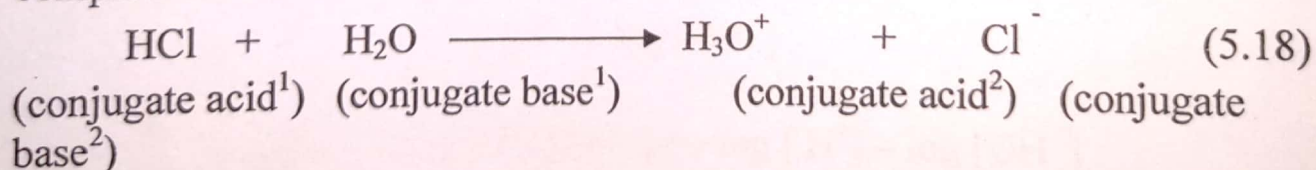
5.3.10 Acid–Base Equilibrium

The Bronsted theory states that an acid is a substance that can donate a proton, and a base is a substance that can accept a proton.

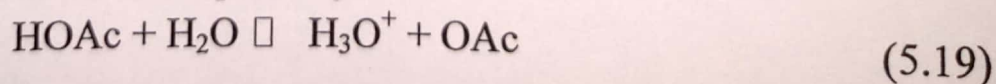


The acid and base pairs in the ionization reaction are called conjugate pairs, that is, the base is a conjugate of an acid and vice versa.

When an acid or a base dissolves in water, it will dissociate or ionize, the amount of ionization being dependent on the strength of the acid. A strong electrolyte is completely dissociated while a weak electrolyte is partially dissociated. Hydrochloric acid is a strong acid and undergoes complete ionization.



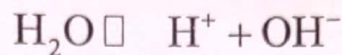
The proton exists in water as the hydronium ion. On the contrary acetic acid is a weak acid and ionizes partially



The equilibrium constant for this reaction may be represented as follows

$$k = \frac{a_{\text{H}_3\text{O}^+} a_{\text{OAc}^-}}{a_{\text{HOAc}} a_{\text{H}_2\text{O}}} \quad (5.20)$$

where k is the acidity constant and a is the activity of the individual species. Calculations are simplified if we neglect activity coefficients and use molar concentration instead. Thus the autoprotolysis of water can be represented as



at 25°C , $k_w = 1.0 \times 10^{-14}$ or $[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$. In pure water the concentration of $[\text{H}^+]$ and $[\text{OH}^-]$ are equal, or $[\text{H}^+] = [\text{OH}^-]$ so $[\text{H}^+][\text{H}^+] = 1.0 \times 10^{-14}$
then $[\text{H}^+] = 1.0 \times 10^{-7} = [\text{OH}^-]$

Box 5.9 Calculate the hydroxyl ion concentration in a 1.0×10^{-2} M solution of hydrochloric acid

Solution $[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$
since $[\text{H}^+] = 1.0 \times 10^{-2}$; $[1.0 \times 10^{-2}][\text{OH}^-] = 1.0 \times 10^{-14}$
or $[\text{OH}^-] = \frac{1 \times 10^{-14}}{1 \times 10^{-2}} = 1 \times 10^{-12} \text{ M}$

5.3.11 The pH Scale

The pH of a solution was defined by Sorensen as the negative logarithm of the hydrogen ion concentration

$$\text{pH} = -\log \text{H}^+ \quad (5.21)$$

A similar definition is made for the hydroxyl ion concentration

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

$$-\log k_w = -\log [\text{H}^+][\text{OH}^-] = -\log [\text{H}^+] - \log [\text{OH}^-]$$

or

$$\text{p}k_w = \text{pH} + \text{pOH}$$

Box 5.10 Calculate the pOH and pH of a solution of NaOH

Solution $[\text{OH}^-] = 3.0 \times 10^{-3} \text{ M}$

$$\text{pOH} = -\log(3.0 \times 10^{-3}) = 3 - \log 3.0 = 3 - 0.477 = 2.52$$

$$\text{pH} + 2.52 = 14.00 \text{ or } \text{pH} = 11.48$$

Box 5.11 The pH of a solution is 8.43. Calculate the hydrogen ion concentration in the solution

Solution $\text{pH} = -\log [\text{H}^+] = 8.43$

$$\text{or } [\text{H}^+] = 10^{-8.43} = 10^{-9} \times 10^{0.57}$$

$$[10^{0.57} = 3.71]$$

$$\text{or } [\text{H}^+] = 3.71 \times 10^{-9}$$