### 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.

$$\text{Acid} = H^+ + \text{base} \tag{5.17}$$

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.

$$HCl + H_2O \longrightarrow H_3O^+ + Cl^- \tag{5.18}$$

(conjugate acid¹) (conjugate base¹) (conjugate acid²) (conjugate base²)

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

$$HOAc + H_2O \rightleftharpoons H_3O^+ + OAc \tag{5.19}$$

The equilibrium constant for this reaction may be represented as follows

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$$k = \frac{aH_3O^+ \ aOAc^-}{aHOAc \ aH_2O} \quad (5.20)$$

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

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

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

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

**Solution** $[H^+]\ [OH^-] = 1.0 \times 10^{-14}$
since $[H^+] = 1.0 \times 10^{-2}$ ; $[1.0 \times 10^{-2}]\ [OH^-] = 1.0 \times 10^{-14}$

$$\text{or } [OH^-] = \frac{1\times10^{-14}}{1\times10^{-2}} = 1\times10^{-12}\text{M}$$

### 5.3.11 The pH Scale

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

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

A similar definition is made for the hydroxyl ion concentration

$$pOH = -\log [OH^-]$$

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

or

$$pkw = pH + pOH$$

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

**Solution** [OH-] = 3.0 × 10-3 M

pOH = - log (3.0 × 10-3) = 3 – log 3.0 = 3 – 0.477 = 2.52

pH + 2.52 = 14.00 or pH = 11.48

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

**Solution** $pH = -\log [H^+] = 8.43$

or $[H^+] = 10\text{-}8.43 = 10^{-9} \times 10^{0.57}$ $[10^{0.57} = 3.71]$

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