E-Content Study Material

B. Sc. Chemistry (H)

2nd Year

Paper II B

Inorganic Chemistry

Chapter VII: Acids and Bases

Topic: Arrhenius and Bronsted-Lowry Concept of Acids and Bases

Prepared By

Dr. Satyendra Gupta

Assistant Professor (Guest Faculty)

Department of Chemistry

Patna Science College, Patna

Email: satyendgupta@gmail.com

Mob. No. 9794346892

Concept of Acids and Bases

Different concepts have been put forth by different investigators to characterise acids and bases. While some of the concepts are quite narrow in their approach, others are fairly comprehensive. Some important concepts of acids and bases are,

Arrhenius Concept

According to Arrhenius concept, an acid is a substance that dissociates to give hydrogen ions when dissolved in water. Thus, hydrogen chloride gas is an acid because when dissolved in water, it gives hydrogen ions.

$$HCl\ (g) + Water \rightleftharpoons H^{+}\ (aq) + Cl^{-}\ (aq)$$

The symbol aq indicates that the ions are hydrated, i.e., associated with one or more molecules of water.

According to Arrhenius concept, a base is a substance which dissociates into hydroxyl ions when dissolved in water. Thus, NaOH is a base because it furnishes hydroxyl ions in aqueous solution.

$$NaOH + water \rightleftharpoons Na^{+} (aq) + OH^{-} (aq)$$

The high dielectric constant of water lowers the forces of attraction between the oppositely charged ions and thus causes the dissociation of the electrolyte.

Hydronium ion. Hydrogen ion is merely a proton. It is a unique ion in the sense that it has no electron. It is the smallest ion known and, therefore, has a strong tendency to get hydrated. It exits largely as attached to a molecule of water forming H_3O^{+} ion, the hydronium ion.

Although the proton in aqueous solution is known to be largely hydrated, with H_3O^{+} as its probable structure, it is customary to represent it as H^{+} (aq), for the sake of convenience.

Due to simplicity, Arrhenius concept was widely accepted and used. However, this concept has a serious limitation that it can be applied only to aqueous solutions. It fails to account for the behaviour of acids and bases in solvents other than water, i.e., non-aqueous solvents.

Bronsted-Lowry Concept (Proton Transfer Theory).

Bronsted and Lowry suggested a more general definition of acids and bases which applies to aqueous as well as to non-aqueous solutions. According to this concept, an acid is defined as a substance which has a tendency to donate a proton to any other substance and a base as a substance which has a tendency to accept a proton from any other substance. In other words, an acid is a proton-donor and a base is a proton-acceptor.

When an acid loses a proton, the residual part of its has a tendency to regain the proton. Therefore, it behaves as a base. An acid and a base may, therefore, be defined by the general equation,

$$\text{Acid} \rightleftharpoons H^+ + \text{Base}$$

Consider ionisation of acetic acid in water which may be represented as

$$\underset{\text{Acid}}{CH_3COOH} + \underset{\text{Base}}{H_2O} \rightleftharpoons \underset{\text{Acid}}{H_3O^+} + \underset{\text{Base}}{CH_3COO^-}$$

It is evident that acetic acid donates a proton to water and thus acts as an acid. Water accepts a proton and, therefore, acts as a base. In the reverse reaction, hydronium ion (H_3O^+) donates a proton to the acetate ion and, therefore, acts as an acid. The acetate ion accepts a proton and, therefore, behaves as a base.

Such pairs of substances which can be formed from one another by the gain or loss of a proton are known as conjugate acid-base pairs. Thus, acetic acid is the conjugate acid of the acetate ion and acetate ion is the conjugate base of acetic acid. Similarly, water is the conjugate base of hydronium ion and hydronium ion is the conjugate acid of water.

The ionisation of hydrochloric acid in water may be represented as,

$$HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$$

Acid Base Acid Base

Evidently, hydrochloric acid is the conjugate acid of chloride ion and chloride ion is the conjugate base of hydrochloric acid.

The following points emerge out of the above discussion:

1. Firstly, it is evident that a substance is able to show its acidic character only if another substance capable of accepting a proton (i.e., a base) is present. For example, acetic acid or hydrogen chloride solution in benzene is not acidic because benzene is not in a position to take up protons. But, a solvent like water can take up protons and, therefore,, acetic acid or hydrochloric acid can ionise in water, as shown in the above examples.
2. Secondly, hydrogen ion in aqueous solution does not exist as H^+ ion but as hydrated ion, H_3O^+.
3. Thirdly, not only molecules but even ions may act as acids or bases. Thus, in the above examples, acetate (or chloride) ion act as a base as it can accept a proton to form acetic acid (or hydrochloric acid). In fact, anion of any acid is the conjugate base of the acid.

As can be seen from the ionisation of acetic acid or hydrochloric acid represented above, there is one acid and one base on each side of the equation. Suppose, acetic acid is designated as $Acid_1$, then the acetate ion, its conjugate base, may be designated as $Base_1$.

Similarly, if water is referred to as $Base_{2,}$ then the hydronium ion, its conjugate acid, may be designated as $Acid_{2.}$ The ionization of acetic acid in water may thus be represented as,

$$Acid_1 + Base_2 \rightleftharpoons Acid_2 + Base_1$$

Acid$_1$ and Base$_1$ is a conjugate acid-base pair and so is the pair of Acid$_2$ and Base$_2$.

In general, the ionisation of an acid HA in water may be represented as

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

Just as an acid requires a solvent that can take up a proton (i.e., can act as a base) for its ionisation, a base requires a solvent that can take up a proton (i.e., can act as an acid) for its ionization. Water possesses basic and acidic properties. Therefore, acids as well as base can ionize in water. Thus, water acts as an acid (a proton0donor) towards ammonia and as a base (a proto-acceptor) towards acetic acid. Such substances are said to be amphiprotic.

Consider ionization of ammonia in water which may be represented as

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$\text{Base}_1 \quad \text{Acid}_2 \quad \text{Acid}_1 \quad \text{Base}_2$$

Now ammonium ion (NH_4^+) is the conjugate acid (say, $Acid_1$) to the base NH_3 ($Base_1$) and hydroxyl ion (OH^-) is the conjugate base ($Base_2$) to the acid H_2O ($Acid_2$).

Relative Strengths of Acid-Base Pairs.

The strength of an acid, according to the above concept of Lowry and Bronsted, depends upon its tendency to lose protons and the strength of a base depends upon its tendency to gain protons. If an acid, such as hydrochloric acid, is a strong acid, it will have a strong tendency to donate protons. Thus, the equilibrium

$$HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$$

Lies very much to the right and the reverse reaction, representing the gain of protons by the chloride ions leading to the reformation of HCl, takes place to a very small extent. Accordingly, Cl- ion is a weaker base than H_2O since the latter has a greater tendency to take up a proton.

Acetic acid, on the other hand, has considerably less tendency to donate protons and is, therefore, a much weaker acid. It ionises to a small extent, i.e., the equilibrium

$$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$$

lies mostly towards the left. It follows, therefore, that CH_3COO^- ion has a much greater tendency to gain protons and hence it is a much stronger base than H_2O.

In the above reaction, acetic acid and hydronium ion may be regarded as two acids competing with each other to donate protons. Again, since the equilibrium is known to lie towards the left, it follows that hydronium ion is more successful in this attempt than acetic acid. Hence, the former is a stronger acid. In short, while acetate ion is a stronger base than water, its conjugate acid, acetic acid, is a weaker acid than hydronium ion.

Thus, as a general rule, the stronger an acid, the weaker must be its conjugate base and vice versa. If an acid (e.g., HCl) is strong, its conjugate base (Cl^- ion) is weak. If a base (e.g., CH_3COO^-) is strong, its conjugate acid (CH_3COOH) is weak.

Water is a weak base because its conjugate acid, hydronium ion (H_3O^+), is a strong acid. At the same time, water is a very weak acid because its conjugate base, hydroxide ion (OH^-), is a very strong base.

It may be pointed out that while strong acids like HCl, HNO_3, H_2SO_4, etc., being covalent compounds, get ionised only in a solvent like water which can take up protons, strong bases such as NaOH, KOH, $Ba(OH)_2$, etc., being electrovalent compounds, exist as ions even in the solid state. The basic character of these compounds is exclusively due to the presence of hydroxyl ion which is always there (even without a solvent) and, therefore, no interaction with a solvent like water is necessary in such cases.