ACIDS, BASES AND SALTS

Chapter 8

Teaching Periods	Assessment	Weightage
10	1	7

Students will be able to:

- **Define** Bronsted and Lowery concepts for acids and bases.
- **Identify** conjugate acid-base pairs of Bronsted-Lowery acid and base.
- **Explain** ionization constant of water.
- **Calculate** pH, pOH in aqueous medium using Kw values.
- **Define** and explain leveling effect.
- **Define** Lewis acid and base with suitable examples.
- **Define** a buffer and make buffer solutions.
- **Show** with equations how a buffer system works.
- **Applications** of salts like NaCl, KCl, KI, $NaHCO_3$, $MgSO_4$, etc.
- **Use** the concept of hydrolysis to explain why aqueous solutions of some salts are acidic or basic.
- **Use** concept of hydrolysis to explain why the solution of a salt is not necessarily neutral.

INTRODUCTION

For centuries, acids and bases have been employed as laboratory chemicals and also used in homes. Acetic acid (CH_3COOH) can be found in our kitchens as vinegar, citric acid ($H_3C_6H_5O_7$) is found in citrus fruits like orange, lemon, etc., and phosphoric acid (H_3PO_4) serves as a flavouring agent in many carbonated beverages. Some common household bases are sodium hydroxide (NaOH) which is used as a drain cleaner, ammonia (NH_3) is used as a glass cleaner and sodium hydrogen carbonate ($NaHCO_3$) serves as baking soda in our cooking items.

An acid is any species that has a sour taste, produces hydrogen gas when it reacts with active metals like aluminium and zinc. On the other hand, a base is any species with a bitter taste and a slippery texture (Avoid tasting or touching laboratory chemicals; instead, use acetic acid in the form of vinegar in salad.) Furthermore, it was well understood that when acids and bases react, each cancels out the other's qualities in a process known as neutralisation.

Fig. 8.1: Common uses of acids, bases and salts

8.1 ACIDIC, BASIC AND AMPHOTERIC SUBSTANCES

Oxides of metals and non metals are not themselves acid or base but when dissolves in water, they form acidic or alkaline solution. On the basis of this fact, oxides are classified into three main groups.

Do You Know?

A wide range of organic acids such as acetic acid, citric acid, ascorbic acid, oxalic acid, tartaric acid and amino acid etc are found in our food. They serve as flavorant, antioxidant, energy producer and microbial inactivitor etc.

(i) Metal oxides such as Na_2O, K_2O, CaO etc readily mixed in water to produce alkaline solution.

$$K_2O_{(s)} + H_2O_{(l)} \longrightarrow 2KOH_{(aq)}$$
$$Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)}$$
$$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)}$$

(ii) Non metal oxides such as CO_2, SO_2, P_2O_5 etc when mixed in water, they produce acidic solution.

$$CO_{2(g)} + H_2O_{(l)} \longrightarrow H_2CO_{3(aq)}$$
$$SO_{2(g)} + H_2O_{(l)} \longrightarrow H_2SO_{3(aq)}$$
$$P_2O_{5(s)} + 3H_2O_{(l)} \longrightarrow 2H_3PO_{4(aq)}$$

(iii) Certain metal oxides such as ZnO, Al_2O_3 etc are not water soluble. However, they have ability to react with both acid and base. These are on the borderline of both acid and base and known as amphoteric substances.

$$Al_2O_{3(s)} + 2NaOH_{(aq)} \longrightarrow 2NaAlO_{2(aq)} + H_2O_{(l)}$$
$$Al_2O_{3(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3(aq)} + 3H_2O_{(l)}$$

Beside oxides, certain other substances are also acidic and basic in nature for example Na_2CO_3 gives alkaline solution in water and NH_4Cl gives acidic solution.

Self Assessment

Identify acidic, basic and amphoteric substance in the following:
SO_2, NH_4Cl, Na_2CO_3, ZnO, Na_2O

8.2 THE BRONSTED-LOWRY THEORY OF ACIDS AND BASES

In 1923 J.N Bronsted (Danish Chemist) and J.M Lowry (British Chemist) individually propose this general approach to explain acids and bases as an extension of Arrhenious concept.


J.N Bronsted J.M Lowry

According to this theory **"Acid is a species which tends to donate proton (protogenic) whereas base is a species that accept proton (protophillic). Further, an acid base reaction is the transfer of proton from acid to base"**.

In the light of this definition all hydrogen containing substances (either molecules or ions) which are capable of giving up proton (HCl, HNO_3, H_2SO_4, H_2O, H_3O^+, NH_4^+etc) are gathered in the list of acids where as all those substances which have ability to accept proton (NH_3, H_2O, OH^-, Cl^-, CO_3^{-2}, HSO_4^-) are categorized as bases.

8.2.1 Proton Donors and Acceptors

According to Bronsted Lowry concept, the necessary need for an acid base reaction is that one substance can donate a proton while other substance is able to accept it. In this proton transfer process both acid and base work together. This means if one species is an acid the other always be a base.

Let us consider what happens when a strong acid (HCl) dissolves in water and undergo ionization.

$$\underset{\text{(acid)}}{HCl_{(l)}} + \underset{\text{(base)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(conjugate acid)}}{H_3O^+{}_{(aq)}} + \underset{\text{(conjugate base)}}{Cl^-{}_{(aq)}} \text{ ------ (i)}$$

Since HCl donate a proton, it acts as an acid on the other hand water (H_2O) accept a proton and serve as base.

Look at the reverse of this acid base reaction where hydroxonium ion (H_3O^+) reacts with chloride ion (Cl^-). In the reverse reaction, H_3O^+ act as conjugate acid while Cl^- ion serves as conjugate base. Thus the equilibrium of the above reaction represents two acids and two bases, one on either side of the reaction arrow.

Now see the ionization of ammonia in to water in the given equilibrium.

$$\underset{\text{(base)}}{NH_{3(g)}} + \underset{\text{(acid)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(conjugate acid)}}{NH^+_{4(aq)}} + \underset{\text{(conjugate base)}}{OH^-_{(aq)}} \text{------ (ii)}$$

Reaction in the forward direction shows that water gives up a proton (acid) and ammonia accept it (base). In reverse phase of this reaction ammonium ion (NH_4^+) acts as conjugate acid and hydroxide ion (OH^-) as conjugate base.

While studying above two acid base reactions it is interesting to note that water serves as base when HCl is dissolved into it but acts as acid when ammonia is added.

8.2.2 Relative Strength of Bronsted Acid and Base

The direction of an acid base reaction depends on the relative strengths of concerned acid and base and the reaction is generally goes towards the direction of weaker acid or base. According to Bronsted-Lowry perception, the strongest acid has the weakest conjugate base and the strongest base has weakest conjugate acid.

Look at again the acid base reaction mentioned in equation (i). Since HCl is a strong acid, it donates proton to water molecules completely and forms a conjugate base of HCl which is Cl^- and conjugate acid of H_2O which is H_3O^+.

$$\underset{\textbf{(Stronger acid)}}{HCl_{(l)}} + H_2O_{(l)} \rightleftharpoons Cl^-_{(aq)} + \underset{\textbf{(Stronger acid)}}{H_3O^+_{(aq)}}$$

In term of relative strength of two acids (HCl and H_3O^+), HCl is stronger than H_3O^+ because it lose proton more readily than H_3O^+.

CH_3COOH on the other hand is a weak acid since it is only 1.4% ionize when dissolve in water therefore it is weaker acid than H_3O^+.

$$\underset{\text{(Weak acid)}}{CH_3COOH_{(l)}} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + \underset{\text{(Strong acid)}}{H_3O^+_{(aq)}}$$

The relative order of the three acids discussed above is now written as

$$HCl > H_3O^+ > CH_3COOH$$

The relative order of strength of bases can be determined in the same way.

$$NaOH > NH_3 > NH_4OH$$

8.3 CONJUGATE ACID BASE PAIRS

In Bronsted-Lowry acid base reaction both reactants and products serves as acid and base.

A conjugate acid base pair represents two species, one on the left side and other on the right side of equation with the difference of one proton. It is important to note that strong acid has weak conjugate base and weak base has strong conjugate acid.

How to identify conjugate acid base pair

Consider the reaction of H_2S and NH_3.

$$H_2S_{(g)} + NH_{3(aq)} \rightleftharpoons HS^-_{(aq)} + NH_4^+{}_{(aq)}$$

To identify a conjugate acid base pair, it should be in our mind that the only difference between the two species is a proton (H^+). Here we see that H_2S change into HS^- by losing a proton. Thus H_2S and HS^- represents a conjugate acid base pair. NH_3 on the other hand accept a proton and change into NH_4^+ therefore it is another conjugate acid base pair.

$$\underset{\text{(Bronsted acid)}}{H_2S} + \underset{\text{(Bronsted base)}}{NH_3} \rightleftharpoons \underset{\text{(Conjugate acid)}}{NH_4^+} + \underset{\text{(Conjugate base)}}{HS^-}$$

Some common acid base pairs are listed below

Bronsted Acid		Bronsted Base	$\rightleftharpoons$	Conjugate Acid		Conjugate Base
HCN	+	H_2O	$\rightleftharpoons$	H_3O^+	+	CN^-
CH_3COOH	+	NH_3	$\rightleftharpoons$	NH_4^+	+	CH_3COO^-
H_2O	+	H_2O	$\rightleftharpoons$	H_3O^+	+	OH^-
HBr	+	F^-	$\rightleftharpoons$	HF	+	Br^-
H_2O	+	CO_3^{-2}	$\rightleftharpoons$	HCO_3^-	+	$O\bar{H}$
HSO_4^-	+	CO_3^{-2}	$\rightleftharpoons$	HCO_3^-	+	SO_4^{-2}

Self Assessment

Identify the acid and base on the right hand side of the following equilibria.

(i) $H_2S + H_2O \rightleftharpoons HS^- + H_3O^+$ (ii) $NH_4^+ + S^{-2} \rightleftharpoons HS^- + NH_3$

Limitations of Bronsted-Lowry Theory

(i) This theory cannot explain the reaction between acidic oxides (CO_2, SO_2, SO_3) and the basic oxides like (CaO, BaO, MgO) in which no involvement of proton transfer.

(ii) Certain substances like $AlCl_3$, BF_3 etc do not involve in the proton transfer but they function as acid.

8.4 STRENGTHS OF ACIDS AND BASES

Consider Bronsted Lowry theory, all acids have ability to donate proton in aqueous medium, however some acids release their proton faster than others. The relative strength of acids and bases can be measured by their degree of ionization in water.

The degree of ionization (α) is the ratio of the number of molecules ionized to the total number of the dissolved molecules. This ratio when multiplied by hundred(100), we can get ionization in percentages.

$$\%\text{age of ionization } (\alpha) = \frac{\text{Number of molecules split into ions}}{\text{Total number of molecules dissolved}} \times 100$$

The %age of ionization of strong acids or bases is up to 90-95%, which means that out of every 100 molecules, 90 to 95 molecules get ionized in aqueous solution. Many acids ionizes only to a limited extent in water, these are called weak acids. The aqueous equilibrium mixture of these acids possess unionized molecules of acid (HA), conjugate base and H_3O^+ ions.

Table 8.1 The percentage ionization of some acids and bases.

Acids	Degree of Ionization	Bases	Degree of Ionization
HCl	90 – 95%	$NaOH$	90 – 95%
HNO_3	90 – 95%	KOH	90 – 95%
H_2SO_4	60% (1st step)	$Ba(OH)_2$	77% (1st step)
CH_3COOH	1.4%	NH_4OH	1.4%

The ionization of an acid in its aqueous solution is a reversible process. We can write the equilibrium reaction of ionization of an acid as under:

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

By applying the law of mass action to the acid ionization equilibrium:

$$Kc = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$

In the dilute solution of an acid (HA) we assume that the concentration of liquid water remains constant.
Therefore

$$Ka = \frac{[H_3O^+][A^-]}{[HA]}$$

The ionization of a base (B) in water, at equilibrium is shown below.

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

Base ionization constant expression may be written in the same way as:

$$Kb = \frac{[BH^+][OH^-]}{[B]}$$

The higher the value of Ka and Kb the stronger is the acid and base.

Table 8.2 Ionization constant values of some weak acids and weak bases at 25°C

Acids	Formula	Ka	Bases	Formula	Kb
Oxalic acid	$H_2C_2O_4$	5.6×10^{-2}	Ethylamine	$C_2H_5NH_2$	4.7×10^{-4}
Formic acid	HCOOH	1.7×10^{-4}	Ammonia	NH_3	1.8×10^{-5}
Acetic acid	CH_3COOH	1.7×10^{-5}	Hydrazine	N_2H_4	1.7×10^{-6}
Carbonic acid	H_2CO_3	4.3×10^{-7}	Hydroxylamine	NH_2OH	1.1×10^{-8}
Boric acid	H_3BO_3	5.9×10^{-10}	Pyridine	C_5H_5N	1.4×10^{-9}

8.4.1 Ionization Constant of Water (Kw)

Although water is chemically non electrolytic substance, however it conducts electricity to a very small extent due to its self ionization ability. During the reaction proton from one water molecule is transferred to another water molecule as given in the following reaction.

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+(aq) + O\bar{H}\,(aq)$$

The equilibrium constant may be written as

$$K = \frac{[H_3O^+][O\bar{H}]}{[H_2O][H_2O]}$$

$K\,[H_2O]^2 = [H_3O^+][O\bar{H}]$ (Hence concentration of water is constant)

$K_W = [H_3O^+][O\bar{H}]$

At 25°C, the value of Kw is 1.0×10^{-14}. This very small value of Kw shows that self ionization of water occurs to a very small extent. Since the number of $[H^+]$ and $[OH^-]$ in pure water are equal, we can determine the concentration of these ions by using above equation of Kw. Thus in pure water the conc. of each $[H^+]$ and $[OH^-]$ at 25°C is 1×10^{-7} M.

The value of Kw varies with temperature and a rise of temperature increases the value of Kw.

8.4.2 pH, pOH and pKw

Soren Sorensen

The measurement of the concentration of aqueous H^+ and OH^- ions is a task to estimate the strength of acids and bases because the conc. of H^+ and OH^- ions are very low in aqueous solution specifically when a weak acid or base is dissolved. To cope up this difficulty Sorenson (1909) introduced a logarithmic scale since logarithm is very useful way in representing very small and very large quantities. **"pH is the negative logarithm of molar concentration of H^+ ions and pOH is the negative logarithm of molar concentration of OH^- ions in aqueous solution at given temperature".** Since pH is a logarithmic value, it has no unit.

$$pH = -\log [H^+] = \log \frac{1}{[H^+]}$$

$$pOH = -\log [OH^-] = \log \frac{1}{[OH^-]}$$

pH scale represents numbers between zero (very acidic) to 14 (very alkaline) and specify the acidity or basicity of aqueous solution. A solution having pH = 7, represents the point of neutrality and it evident that solution contains equal conc. of H^+ ions and OH^- ions. The behavior of different solutions with respect to pH concept is summarized below.

Table 8.4 PH of acidic, basic and neutral solution at 25°C

Aqueous Solution	Conc. of H^+ and OH^- ions	pH
Neutral	$[H^+] = [OH^-]$	7
Acidic	$[H^+] > [OH^-]$	Less than 7
Alkaline	$[H^+] < [OH^-]$	More than 7

Although, the pH of distilled water is 7 at 25°C, however it increases with the rise of temperature since temperature increases the extent of ionization.

The dissociation constant of water is represented as

$K_W = [H^+] [OH^-]$

By taking negative logarithm on both sides

$(-\log K_W) = (-\log [H^+]) + (-\log [OH^-])$

Since K_W at 25°C is 1×10^{-14}

Therefore

pH + pOH =14

Do You Know?

Honey bee when stings a person, it injects a painful and irritating acid (formic acid) into the skin. It can give a relief if a mild base like backing soda is rubbed on the stung area of skin.

pH of a solution is mostly determined by using universal indicator. It is mixture of dyes usually pasted on a paper. It shows a change in the colour with the change of PH and helps us for measuring the strength of acids and bases. Several other indicators are known and each undergoes a colour change over a particular pH range.

Table 8.5 Colour of some indicators in acid, base and neutral conditions.

Indicator	Colour in acid	Colour in Base	Colour when Neutral
Litmus	Red	Blue	Purple
Methyl orange	Red	Yellow	Orange
Phenolphthalein	Colourless	Pink	Colourless
Universal Indicator	Red	Purple	Green

pH play a vital role in many biological and industrial process. Following is the list of pH of some common substances at 25°C.

Table 8.6 pH values of some common known substances

Items	pH	Items	pH
Urine	6.0	Tomato	4-4.1
Human blood	7.35-7.45	Lemon juice	2.3
Milk	6.6	Soft drinks	3
Sea water	7.9-8.4	Milk of magnesia	10.5
Pure water	7.0	Stomach	1.6-1.7
Vinegar	2.8-3	Egg white	8.0-8.1
Orange juice	3.5-3.7	Wash soda	12.0

Example 8.1

Calculate the pH of 0.0025M aqueous solution of HCl at 25°C

Solution:

Since HCl is a strong acid, it is assumed to be completely ionized in aqueous solution.

$$\underset{(0.0025M)}{HCl_{(l)}} \rightleftharpoons \underset{(0.0025M)}{H^+_{(aq)}} + Cl^-_{(aq)}$$

Therefore for every molecule of HCl, there is one H^+ ion, hence

$$[H^+] = [HCl]$$

or $$[H^+] = 0.0025M$$

Now using the formula of pH

$$pH = -\log [H^+]$$
$$pH = -\log [0.0025]$$
$$pH = 2.6$$

Example 8.2

Determine the pH of 0.1M NaOH solution at 25°C.

Solution:

NaOH is a strong base and it is assumed to be completely ionized in aqueous solution.

$$\underset{(0.1M)}{NaOH} \rightleftharpoons Na^+ + \underset{(0.1M)}{OH^-}$$

Therefore the concentration of hydroxide ions (OH^-) is equal to the concentration of dissociated NaOH.

$$[OH^-] = [NaOH] = 0.1M$$

Now applying the relation

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

$pOH = -\log [0.1]$

$pOH = 1$

Since $pH + pOH = 14$

$pH = 14 - 1$

$pH = 13$

Example 8.3

The pH of a solution 25°C is 4.82, calculate its hydrogen ion concentration.

Solution:

Since $pH = -\log [H^+]$

Taking antilog of this equation, we get the following relationship

$[H^+] = 10^{-PH}$

$[H^+] = 10^{-4.82}$ (Take antilog of $10^{-4.82}$)

$[H^+] = 1.51 \times 10^{-5} M$

Self Assessment

The pH of a sample of urine is noted to be 5.3. Calculate $[H^+]$ and $[OH^-]$ ions in the solution.

8.4.3 Leveling Effect

Strong acids like HCl, HBr, HI etc when dissolve in water, they all exhibit apparently equal strength due to the fact that water play basic medium to get them ionized completely hence water serves as leveling solvent for these acids and the phenomena is called leveling effect.

Therefore, any acid stronger than hydroxonium ion (H_3O^+) completely reacts with water to form H_3O^+ ion and the corresponding conjugate base. This can be attributed to the fact that H_3O^+ is the strongest acid that can possibly exist in an aqueous solution.

For example HCl completely react with water to form H_3O^+ and Cl^-.

$$HCl_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$$

The similar concept goes for bases as well. Any base stronger than hydroxide ion (OH^-) completely react with water to form OH^- ion and respective conjugate acid. The reason is similar to our former understanding that OH^- ion is strongest base that can possibly exist in any aqueous solution.
For example when sodium amide ($NaNH_2$) is placed in water then amide ion (NH_2^-) reacts completely with H_2O.

$$NH^-_{2(g)} + H_2O_{(l)} \longrightarrow NH_{3(aq)} + OH^-_{(aq)}$$

8.5 LEWIS DEFINITION OF ACIDS AND BASES

According to G.N Lewis (1923) **"Acid is a substance that accepts an electron pair whereas base is a substance that can donates electron pair"**. Lewis acid and base is commonly known as electrophillic (electron loving) and nucleophillic (nucleus loving) reagents since they have capability to accept and donate electron pair respectively.

Considering Lewis concept, hydrogen ion (H^+) is an acid because it accepts electron pair. Similarly hydroxide ion (OH^-) is base since it donate lone pair of electron.

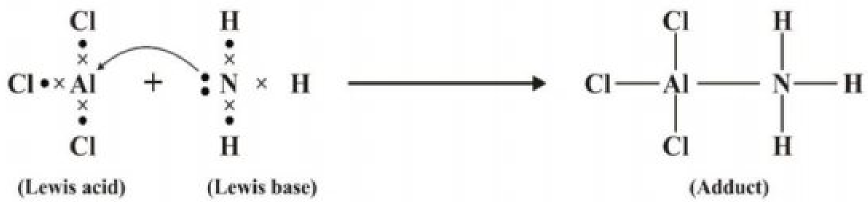
Do You Know?

Lewis acids are not only limited to $AlCl_3$ and BF_3 type molecules but molecules with polar multiple bonds as well as metal cations also function as lewis acids.

Lewis theory does not explain about the transfer of proton from one specie to another however it indicates sharing of electron pair between a donor and accepter reagent. Generally species (compounds or cations) having less than a full octet of electrons behave like Lewis acids and all species (compounds or anions), having lone pair of electrons behave like Lewis base. The product formed in a lewis acid base reaction possesses a coordinate covalent bond between lewis acid and base and known as adduct.

$$A + :B \longrightarrow A-B$$

Several neutral molecules such as $AlCl_3$, BF_3, $FeCl_3$ etc serves as Lewis acids due to electron deficiency on the central atom (fewer than eight valence electrons). On the other hand ammonia (NH_3), Phosphene (PH_3), Water (H_2O) etc although have their complete octet but possess lone pair of electron that is available to donate hence they serves as Lewis base.



8.6 BUFFER SOLUTIONS AND THEIR APPLICATIONS

It is often necessary to maintain the pH of certain solutions in the laboratory and industrial processes. This can be achieved by the help of buffer solutions. **"A buffer solution is one whose pH is not changed significantly on dilution or even if the small amount of acid or base is added at constant temperature".** Buffer solution is a mixture of a weak acid and its conjugate base, or a weak base and its conjugate acid. It resists the change in pH and can keep it for long time.

There are two types of buffer solution.

Do You Know?

Buffering in blood is crucial to our survival. The pH of blood must be kept constant for normal body functions to work. If blood becomes too acidic, or too basic, then enzymes and proteins are unable to function.

Acidic Buffers: It contains weak acid and its salt with strong base, its PH is less than 7. For example: CH_3COOH and CH_3COONa.

Basic Buffers: It contains weak base and its salt with strong acid, its PH is more than 7. For example: NH_4OH and NH_4Cl.

Buffer Action

To understand how a buffer works, let us consider the ionization of acetic acid-sodium acetate solution.

$$CH_3COOH_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$$

$$CH_3COONa_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + Na^+_{(aq)}$$

Both acetic acid and sodium acetate provides CH_3COO^- ions, however CH_3COO^- ions comes from CH_3COONa (strong basic salt) are in high concentration.

When small amount of an acid is added to this buffer solution, the most of the H^+ ions of acid combine with excess acetate ions (CH_3COO^-) of the solution to form acetic acid.

$$CH_3COO^-_{(aq)} + H^+_{(aq)} \longrightarrow CH_3COOH_{(l)}$$

Thus a very slightly change in the PH is observed and we say that PH remains practically unchanged.

When small amount of a base in added to this buffer solution the additional OH^- ions combine with H^+ ions of the buffer to form water molecules. As a result the equilibrium shifts to the right to produce more H^+ ions till practically all the excess of OH^- ions are neutralized and the original buffer pH is restored.

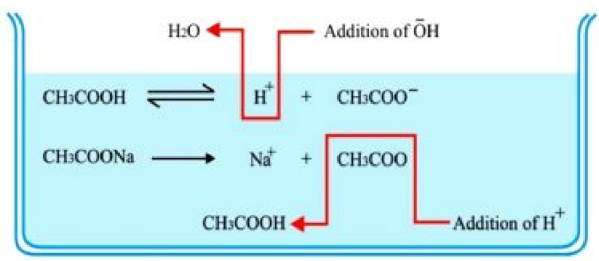


Fig 8.2: Buffer action of an acidic buffer

Buffer Capacity

It is important to realize that buffer solution cannot keep the PH constant if enough acid or base is added into it. The quantity of substance (acid or base) that can be mixed into buffer solution before its PH change is known as buffer capacity.

Application of Buffer solutions:

1. Buffer solution plays a very significant role in biochemical system. For example pH of our blood is maintained at 7.3 to 7.4 due to bicarbonate and carbonic acid buffer.
2. Buffer solutions are widely used in industrial processes such as fermentation, dye processes and manufacturing of pharmaceuticals.
3. Buffer solution is used in agriculture to maintain the pH of soil for proper crop yield.
4. Buffer solution is extensively used in analytical chemistry and pathological laboratories.
5. Buffers are also used in foods industries to maintain the pH of various food items in order to preserve their flavor, appearance and micro-biological stability.

8.7 SALTS THEIR TYPES AND APPLICATIONS

"Salt is a substance produce by the reaction of acid and base. It consists of positive ion of base and negative ion of acid".

Salts may be neutral, acidic or basic, depending upon the number and types of ions present.

Table 8.7 Types of salts formed by neutralization reaction

Acid	Base	Types of Salts	Neutralization Reaction
Strong	Strong	Neutral	$NaOH + HCl \rightarrow NaCl + H_2O$
Strong	Weak	Acidic	$HCl + NH_4OH \rightarrow NH_4Cl + H_2O$
Weak	Strong	Basic	$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$
Weak	Weak	Neutral	$CH_3COOH + NH_4OH \rightarrow CH_3COONH_4 + H_2O$

Hydrolysis

When a salt (AB) dissolves in water, it breaks up into A^+ and B^- ions. These positive and negative ions may or may not react with water. If any one of these ions reacts with water, the solution becomes either acidic or alkaline.

$$A^+ + H_2O \longrightarrow AOH + H^+$$
$$B^- + H_2O \longrightarrow BH + \overline{OH}$$

Thus, the reaction of cation or anion of the salt with water to produce hydrogen ions (H^+) or hydroxyl ions (OH^-) and change the pH of solution is known as hydrolysis (hydro; water and lysis; to break).

8.7.1 Salt and types of Salt

After getting the concept of hydrolysis we shift towards four possible types of salts.

Type-1: (Salt of Strong acid and Strong base)

These salts are formed by neutralization of strong acid and strong base. Since both cation and anion comes from strong base and strong acid respectively, they do not undergo hydrolysis and hence the pH of their aqueous solution remains 7.
For example: NaCl, K_2SO_4, $NaNO_3$ etc

Type-2: (Salt of Strong acid and Weak base)

These salts are formed by neutralization of strong acid and weak base. The an ion of these salts comes from strong acid and hence does not react with water, however cat ion which comes from weak base interact with water to give acid solution having pH less than 7.
For example: When NH_4Cl dissolves in water, it gives acidic solution

$$NH_4Cl_{(s)} \rightleftharpoons NH_4^+ + Cl^-_{(aq)}$$

$$NH_4^+ + H_2O \rightleftharpoons NH_4OH + H^+$$

Other examples of this type of salts are $(NH_4)_2SO_4$, $CuCl_2$ etc

Type-3: (Salt of weak acid and strong base)

These salts are formed by neutralization of weak acid and strong base. The cat ion which comes from strong base does not react further with water. The an ion which comes from weak acid reacts with water and undergo hydrolysis to give alkaline solution having pH more than 7. For example sodium carbonate when dissolve in water it gives basic solution.

Do You Know?

The digestion of food is an example of hydrolysis. The water helps to break down the compounds you have eaten. This allows the larger compounds to be broken down into smaller compounds, so they are more easily absorbed.

$$Na_2CO_{3(s)} \rightleftharpoons 2Na^+_{(aq)} + CO^{-2}_{3\,(aq)}$$

$$CO_3^{-2} + 2H_2O \rightleftharpoons H_2CO_3 + 2OH^-$$

Other examples of this type of salt are CH_3COONa, $NaHCO_3$ etc

Type-4: (Salt of Weak acid and weak base)

These salts are formed by neutralization of weak acid and weak base. Cation and anion of these salts come from weak base and weak acid respectively therefore they both have ability to undergo hydrolysis in aqueous solution. pH of aqueous solution of these salts may either be more than 7, less than 7 or exactly 7 depending upon the relative extent of hydrolysis of cat ion and anion.
For example: NH_4CN, CH_3COONH_4

8.7.2 Application of some salts

Salts play significant role in the synthesis of a variety of industrial products as well as many of them directly use in our daily life. Some applications of salts are mentioned in the following table.

Table 8.8 Applications of some typical salts

Chemical Name	Common Name	Formula	Common Applications
Table salt	Sodium chloride	$NaCl$	Essential part of Human diet
Epsom salt	Magnesium sulphate	$MgSO_4$	Use to produce laxative effect to treat constipation
Gypsom salt	Calcium sulphate	$CaSO_4$	In making plaster of paris
Baking soda	Sodium bicarbonate	$NaHCO_3$	Use as antacid to relieve heart burn and acid indigestion
Washing soda	Sodium carbonate	Na_2CO_3	Use to remove rigid stains from Laundry
Salammoniac	Ammonium chloride	NH_4Cl	Use as expectorant in cough syrup
Muriate of potash	Potassium chloride	KCl	Use in making fertilizers

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Curdling of Milk

Fresh milk is an example of a colloid. It consists of fats and protein particles floating in water. This colloidal suspension scatters light and causes milk to appear white. The protein molecules (mainly casein) repel each other due to the presence of negative charges on them and help them naturally distribute evenly throughout liquid.

Milk is slightly acidic, and its pH is further lowered by milk bacteria which converts lactose into lactic acid or by the addition of some other acidic ingredients like lemon juice, vinegar etc. The protons (H^+) of these acids attach with negatively charged colloidal particles of protein and neutralize them. They now stop repelling each other. This allows them to stick together and coagulate into the clumps known as curds. The watery liquid left behind is called "Whey".

Activity

This activity will enable you to understand how an indicator works in acidic and basic medium.

Place a few small pieces of red cabbage in a bowl of boiling water, stir well, and let aside for 5 minutes, or until a dark solution forms. Now filter it and divide it between two beakers. In one beaker, place the egg white, and in another, the vinegar. It will be observed that the cabbage juice acquires green color in egg white and red in vinegar exhibiting its basic and acidic nature respectively. This is due to the presence of anthrocyanin pigment in cabbage, which causes the colour to change and so acts as an indicator.

SUMMARY with Key Terms

- **Bronsted-Lowry theory** tells us that acid is a specie which tends to donate proton where as base is a specie that accept proton. It also tells that strong acid has weak conjugate base and strong base has weak conjugate acid.
- **Conjugate acid base pair** represents two species, one on the left side and other on the right side of equation with the difference of one proton.
- **Ionization constant of water** is the ionic product of H^+ ion and OH^- ion of water. Its value is 1×10^{-14} at 25°C.
- **pH Scale** is a logarithmic scale use to estimate the strength of acids and bases. It represents number between 0 (very acidic) to 14 (very basic).
- **Universal Indicator** is a mixture of dyes usually paste on a paper. It shows the change in colour with the change of pH and helps us for measuring the strength of acid and base.
- **Leveling effect** is the effect of solvent on the properties of acid and base. All acids stronger than H_3O^+ are completely reacts with water to form H_3O^+ ion and the corresponding conjugate base.
- **Lewis theory of acid and base** tells us that acid is a substance that can accepts an electron pair where as base is a substance that can donates electron pair.
- **Buffer Solution** is a mixture of weak acid with its strong basic salt or a weak base with its strong acidic salt. It tends to resist the change of PH.
- **Buffer Capacity** is the quantity of acid or base added to buffer solution before changing its pH. Buffer capacity is maximum when acid to salt ratio or base to salt ratio is equal to 1.
- **Buffer Range** is the range of pH over which a buffer solution remains effective on addition of strong acid or base.
- **Acid Buffers** are solutions which contain large amount of a weak acid and its salt with strong base. Their pH is less than 7.
- **Basic Buffers** are solutions having large amount of weak base and its salt with strong acid. Their pH is more than 7.
- **Degree of Ionization** is the ratio of number of molecule ionized and the total number of molecules of an acid or base in aqueous solution.
- **Hydrolysis** is the reaction of cation or anion of the salt with water molecule to change the pH of solution cation hydrolysis gives acid solution where as an ion hydrolysis give basic solution.
- **Salts** are formed in the neutralization reaction of acids and bases. These are classified into acidic basic and neutral salts depending upon the number and the type of radicals present in them.

EXERCISE

Multiple Choice Questions

1. Choose the correct answer

(i) H_2SO_4 is stronger acid than CH_3COOH because:
(a) It gives two H^+ ion per molecule (b) Its boiling point is high
(c) Its degree of ionization is high (d) It is highly corrosive

(ii) Al_2O_3 is amphoteric oxide because it reacts with:
(a) Acids (b) Base
(c) Both acids and base (d) neither acid nor base

(iii) Which of the following is not a Buffer solution:
(a) CH_3COOH/CH_3COONa (b) $Na_2CO_3/NaHCO_3$
(c) NH_4OH/NH_4Cl (d) $NaOH/HCl$

(iv) Which oxide is amphoteric in nature:
(a) K_2O (b) CO_2
(c) CaO (d) Al_2O_3

(v) Which of the following does not alter the pH of a solution?
(a) NH_4Cl (b) Na_2CO_3
(c) $NaCl$ (d) $Mg(OH)Cl$

(vi) Conjugated acid of NH_3 is:
(a) NH_4^+ (b) NH_2^-
(c) NH_2 (d) NH

(vii) Salt which is formed by the neutralization of weak acid and strong base is:
(a) $NaNO_3$ (b) NH_4Cl
(c) Na_2CO_3 (d) NH_4CN

(viii) Which of the following statements is not correct about the bases?
(a) They have bitter tastes (b) They have high pH value
(c)They react with acids to form salts (d) They turn blue litmus red

(ix) A conjugate acid base pair has the difference of only:
(a) One electron (b) One proton
(c) One electron pair (d) One proton pair

(x) Salt formed by neutralization of weak acid and weak base is:
(a) NH_4Cl (b) Na_2CO_3
(c) NH_4CN (d) Na_2SO_4

Short Questions

1. Define pH & pOH of a solution? Also show that pH + pOH = 14.
2. Why the aqueous solution of NH_4Cl is acidic and Na_2CO_3 is alkaline.
3. Write down conjugate base of each of the following acids.
   H_2SO_4, H_2S, NH_4^+, HCOOH
4. What is meant by self-ionization of water? Write the expression of Kw. What is its value at 25°C?
5. Write equation and indicate the conjugate acid-base pairs for the following:
   (i) Acetic acid & ammonia (ii) Hydrochloric acid & water

Descriptive Questions

1. Explain Bronsted-Lowry theory of acids and bases. What is meant by conjugate acid-base pair give examples?
2. Define the process of hydrolysis. Explain the behavior of each of the following salts in aqueous solution. (i) K_2CO_3 (ii) $(NH_4)_2SO_4$ (iii) $NaNO_3$
3. What is Buffer solution? Explain how it resists the change of pH by adding small amount of acid and base. Give the applications of buffer solution.
4. Describe Lewis theory of acids and bases. What are the advantages of this theory over Lowry Bronsted theory?
5. What is salt? Explain Acidic, Basic and Neutral salts?

Numerical Questions

1. A solution is made by dissolving 14.8g HCl in water at 25°C. If the volume of solution is 750cm^3 and HCl is assumed to be completely ionized, calculate its pH.
   [Ans: pH = 1.2]
2. 2. The hydroxide ion concentration in an antiseptic solution at 25°C is 3.5×10^{-4}M. Calculate its pH. [Ans: pH = 8]