8

ACIDS, BASES AND SALTS

This is a 8 days unit

After completing this lesson, you will be able to:

- Define Bronsted and Lowery concepts for acids and bases.
- Define salts, conjugate acids and conjugate bases.
- Make a buffered solution and explain how such a solution maintains a constant pH, even with the addition of small amounts of strong acid or strong base.
- Use concept of hydrolysis to explain why the solution of a salt is not necessarily neutral.
- Define and explain levelling effect.
- Explain ionization constant of water and calculate pH and pOH in aqueous medium using given Kw values.
- Use the extent of ionization and the acid dissociation constant, Ka, to distinguish between strong and weak acids.
- Use the extent of ionization and the base dissociation constant, Kb, to distinguish between strong and weak bases.
- Define a buffer, and show with equations how a buffer system works.
- Use the concept of hydrolysis to explain why aqueous solutions of some salts are acidic or basic.
- Identify conjugate acid-base pairs of Bronsted-Lowery acid and base.

Reading

INTRODUCTION

Acids were first recognized as substances that taste sour. Vinegar tastes sour because it is a dilute solution of acetic acid. Citric acid is responsible for the sour taste of a lemon. Bases, sometimes called alkalies are characterized by their bitter taste and slippery feel. Commercial preparations for unclogging drains are highly basic. A salt is an ionic substance that results from the neutralization of an acid and a base. Acid-base chemistry is important in a wide variety of everyday applications. The influence of acids on living things has assumed special importance in recent years due to the phenomenon of acid rain.

The acids classified into (i) Mineral acids or (ii) organic acids. The organic acids are much weaker than mineral acids. The organic acids are mostly found in vegetables, fruit and other stuffs. Some common organic acids and their appearance in different categories are as follows:

Organic acids	Where it is found
Lactic acid	Sour Milk
Citric acid	Citrus fruits like lemons, oranges
Formic acid	Insect bites
Tartaric acid	Grape juice
Maleic acid	Apples and

The presence of water is essential for the formation of H^+, for example HCl is covalent in nature and does not form H^+ ions. However it forms H_3O^+ ions in the presence of H_2O.

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

Bases form a class of chemical substances including metal oxides and hydroxides. A soluble base is called an alkali and forms OH^{-1} ions when dissolved in H_2O. In general the bases on hydrolysis produce alkali.

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Here CaO is basic in nature where as $Ca(OH)_2$ is an alkali and a base.

A salt may be formed by the neutralisation of an acid and a base. A salt may be neutral, acidic or basic

Alkalies are important in soap and detergent manufacture. Caustic soda (NaOH) is used for this purpose.

8.1 ACIDIC, BASIC AND AMPHOTERIC SUBSTANCES

The characteristics of acids and bases are well known e.g. acids turn blue litmus red and react with carbonates to evolve carbon dioxide and bases turn red litmus blue. Hydrochloric acid and nitric acid are common examples of acids, and sodium hydroxide and potassium hydroxide are common examples of bases. However, there are certain substances which are not acids or bases themselves but show acidic or basic nature when dissolved in water. For example, there are certain oxides of metals, like sodium oxide (Na_2O) and calcium oxide (CaO) which react with water to furnish bases sodium hydroxide (NaOH) and calcium hydroxide [$Ca(OH)_2$].

Do You Know

Orious release a gas which turns into sulphuric acid when it reacher you eyes, making them burn.

$$Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)}$$

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

Thus, these oxides are basic in nature. On the other hand certain non-metal oxides like carbon dioxide (CO_2) sulphur dioxide (SO_2) and nitrogen pentaoxide (N_2O_5) when react with water yield carbonic acid (H_2CO_3) sulphurous acid (H_2SO_3) and nitric acid (HNO_3). Such oxides are thus acidic in nature.

Do You Know

Lemon juice contains citric and ascorbic acid.

$$CO_{2(g)} + H_2O_{(l)} \longrightarrow H_2CO_{3(aq)}$$

$$SO_{2(g)} + H_2O_{(l)} \longrightarrow H_2SO_{3(aq)}$$

$$N_2O_{5(g)} + H_2O_{(l)} \longrightarrow 2HNO_{3(aq)}$$

The student should be guided to recall or reproduce information.

Certain oxides are on the border line of being acidic or basic. These oxides which tend to be insoluble in water, are soluble in both acids and bases. They are said to be amphoteric in character that is both acidic and basic. For example aluminium oxide (Al_2O_3) is amphoteric and reacts with both acidic and basic solutions:

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

$$\text{(acidic) } Al_2O_{3(s)} + 2NaOH_{(aq)} \longrightarrow 2NaAlO_{2(aq)} + H_2O_{(l)}$$

Also, chromium trioxide is amphoteric since it reacts with both acids and bases.

$$\text{(basic) } Cr_2O_{3(s)} + 6HNO_{3\,(aq)} \longrightarrow 2Cr(NO_3)_{3\,(aq)} + 3H_2O_{(l)}$$

$$\text{(acidic) } Cr_2O_{3\,(s)} + 6KOH_{(aq)} \longrightarrow 2K_3CrO_{3(aq)} + 3H_2O_{(l)}$$

In addition to oxides, certain salts are also acidic and basic in nature e.g. ammonium chloride (NH_4Cl) is acidic because when dissolved in water it furnishes an acidic solution; potassium carbonate (K_2CO_3) is basic since it yields a basic solution in water.

$$NH_4Cl + \overset{H_2O}{HOH} \longrightarrow NH_4OH + HCl\,.$$

8.2 BRONSTED LOWERY CONCEPTS FOR ACIDS AND BASES:

The concept of acids and bases was further elaborated by J. Bronsted and T. Lowry, independently, by defining them in a different manner. According to these definitions, a Bronsted acid is a proton donor and Bronsted base is a proton acceptor. In this respect, a proton is defined as the nucleus of a hydrogen atom (H^+) and it has nothing to do with the protons in a carbon atom, or a sodium atom or any other atom. Hydrochloric acid is a Bronsted acid since it donates a proton to water according to the following equation.

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

Acid Base

Since water accepts a proton, it is a Bronsted base. After accepting a proton, water is converted to hydroxonium ion (H_3O^+). The hydroxonium ion is infact a hydrated proton. Some more common Bronsted acids are sulphuric acid (H_2SO_4), phosphoric acid (H_3PO_4) and acetic acid (CH_3CO_2H) etc.

The Bronsted bases have been defined as those species which accept a proton. Thus, in the following equation, water is a Bronsted base since it accepts a proton from sulphuric acid.

$$H_2SO_{4(l)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + HSO^-_{4(aq)}$$

Acid Base

Also, in the equation below, ammonia accepts a proton from water. Therefore, ammonia is a Bronsted base and water is a Bronsted acid here:

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

Base Acid

It is important to note that water behaves as a Bronsted acid in presence of ammonia and as a Bronsted base in presence of hydrochloric acid or sulphuric acid. There are many other examples of this type of behaviour. For example, in the equations given below acetic acid is Bronsted acid or a Bronsted base in the second equation.

Interesting Information

The wild lupin plant takes nitrogen from the atmosphere and produces ammonia. It was that ammonia to fertilize the soil for itself and surrounding plants.

$$CH_3COOH_{(l)} + H_2O_{(l)} \longrightarrow CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

Acid Base

$$CH_3COOH_{(l)} + HCl_{(aq)} \longrightarrow CH_3COOH^+_{2(aq)} + Cl^-_{(aq)}$$

Base Acid

8.2.1 Salt, Conjugated Acids and Conjugated Bases

Salts: The joinc substance obtained as a result of neutralization of acids and bases are called salts.

e.g. $HCl + NaOH \longrightarrow NaCl + H_2O$

which substance is a salt in this reaction

Conjugate acids and conjugate bases:

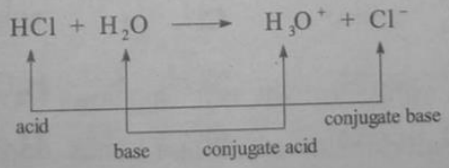
Conjugate acid is a species which is formed as a result of acceptance of proton by a base. Every Bronsted acid has a conjugate base.

Do You Know

A strong acid produces a relatively weak conjugate base. Likewise a strong base. Likewise a strong base produces a relatively weak conjugate acid.

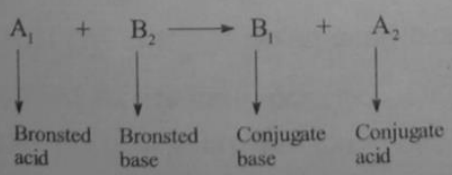
Conjugate base is a species which is left behind after donation of a proton from the acid.

e.g. Ionisation of HCl in water

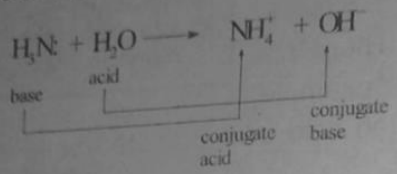


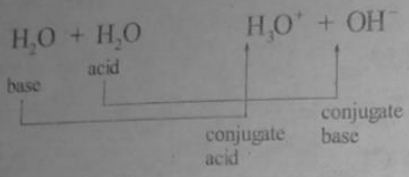
8.3 CONJUGATE ACID – BASE PAIRS

In an acid-base reaction, an acid yields a base (conjugate) and base after accepting proton yields a conjugate acid. The acid-base reaction is represented as



The conjugate acid-base pairs are species on opposite sides of an equation that differ by a proton. The weaker acids have strong conjugate base pairs and stronger acids have weaker conjugate bases e.g.

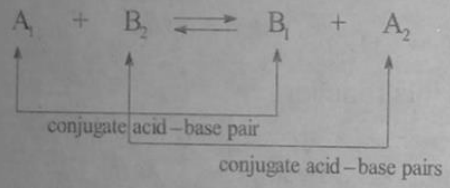




The water is amphoteric in nature i.e. it is acidic as well as basic in nature.

8.3.1 Identification of Conjugate Acid–Base Pairs

The general notation of conjugate acid-base pair is



Each acid has a conjugate base and each base has a conjugate acid.

Particular Examples of Conjugate Acid-Base Pairs:

A_1	+	B_2	$\rightleftharpoons$	B_1	+	A_2
Acid HCl	+	Base H_2O	$\rightleftharpoons$	Cl^{-1}	+	H_3O^{+1}
Acid HCl	+	NH_3	$\rightleftharpoons$	Cl^{-1}	+	NH_4^{+1}
HNO_3	+	H_2O	$\rightleftharpoons$	NO_3^{-1}	+	H_3O^{+1}
NH_4^{+1}	+	H_2O	$\rightleftharpoons$	NH_3	+	H_3O^{+1}
Acid H_2O	+	Base H_2O	$\rightleftharpoons$	OH^{-1}	+	H_3O^{+1}
H_2O	+	CO_3^{-2}	$\rightleftharpoons$	OH^{-1}	+	HCO_3^{-1}
$HClO_4$	+	H_2O	$\rightleftharpoons$	ClO_4^{-1}	+	H_3O^{+1}

A_1 B_1 and B_2 A_2 are known as conjugate acid-base pairs.

8.4 STRENGTH OF ACIDS AND BASES:

Different Bronsted acids donate proton to different extents. An acid which can donate proton to a higher degree than another acid is said to be relatively strong acid. For example, hydrochloric acid is a relatively stronger acid than acetic acid. Also, acetic acid is relatively stronger than water. The ability of an acid to donate proton is called 'strength of acid' or the 'acid strength'. Similarly, .the bases also differ in their ability to accept proton. A base which can accept proton to higher degree than another base is a relatively stronger base. Thus ammonia is a relatively stronger base than water because ammonia can accept a proton to a higher degree than water.

8.4.1 Ionisation constant of Water and Calculation of pH and pOH in Aqueous Medium using Given Kw Values

Water is a unique compound due to its ability to accept or donate proton under different environments. It has been mentioned earlier that water acts as a Bronsted acid in presence of ammonia, and as a Bronsted base in presence of hydrochloric acid. In fact, Water itself undergoes ionization to a small extent as shown in the following equation:

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^- \quad [\ ]$$

This reaction is regarded as auto ionization of water. Here, a molecule of water (which is acting as an acid) donates a proton to another molecule of water (which is acting as a-base) and two species, a hydronium ion and a hydroxide ion, are formed. In this equation, there arc two acid-base conjugate pairs; the first one is water-hydroxide ion (acid-conjugate base) and the second water hydronium ion (base conjugate acid). In fact, there is an equilibrium between water molecules (on the left side of the equation) and the hydronium ions and hydroxide ions on the right side of the equation). The equilibrium constant (K) for this equilibrium can be expressed by the following equation.

$$K = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

In this equation $[H_3O^+]$ $[OH^-]$ and $[H_2O]$ arc the concentration of hydronium ion, hydroxide ion and water at equilibrium stage. Since water is a solvent, the concentration of water, $[H_2O]$ is in large excess, therefore it remains constant. The above equation may be re-written as:

$$K[H_2O]^2 = [H_3O^+][OH^-]$$

the term $K[H_2O]^2$ is the product of two constants, and is

$$K[H_2O]^2 = K_w = [H_3O^+][OH^-] \quad \therefore H_3O^+ = H^+$$

K_w is termed as the ion-product constant of water. Since $[H_3O^+]$ is the concentration of hydrated protons at equilibrium, the above equation corresponds to:

The Student must be guided to use this information in various ways.

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

In pure water at 25°C, the concentration of H^+ and OH^- are equal and found to be 1.0×10^{-7} M each. Thus:

$$K_w = [1.0 \times 10^{-7}][1.0 \times 10^{-7}] = 1.0 \times 10^{-14}$$

It has been noted that whether in pure water or in a solution of dissolved species, the following relationship always holds.

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

Whenever $[H^+] = [OH^-]$, the aqueous solution is found to be neutral, neither acidic nor basic.

$[H^+] > [OH^-]$ Acidic solution

$[OH^-] > [H^+]$ Basic solution

It is to be noted that, because K_w is an equilibrium constant, it is temperature dependent thus, at 40 °C $K_w = 3.8 \times 10^{-14}$ which corresponds to

$$[H^{+1}] = 1.9 \times 10^{-7} \text{ M and } [OH^{-1}] = 1.9 \times 10^{-7} \text{ M as } [H^+] = [OH^-]$$

Since the concentration of $[H^+]$ and $[OH^-]$ are usually very small numbers and inconvenient to work with, a more practical measure called pH was proposed and defined as:

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

It means that pH of a solution is given by the negative logarithm of the $[H^+]$ concentration (in mol/dm^3). However it must be kept in mind that pH being a logarithmic value, does not have any units, the pH concept implies that at 25°C, the different types of solutions will show the following behaviours:

Acid solution: $[H^+] > 1.0 \times 10^{-7}$ M, pH < 7.00

Basic solution: $[H^+] < 1.0 \times 10^{-7}$ M, pH > 7.00

Neutral solution: $[H^+] = 1.0 \times 10^{-7}$ M, pH = 7.00

A scale analogous to the pH can be devised using the negative logarithm of the H^+ concentration.

Example 8.1:

The concentration of $[OH^-]$ ion in a household ammonia solution is 0.005M. Calculate the concentration of $[H^+]$ in it.

Solution:

$[OH^-] = 0.005$ M

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

$1.0 \times 10^{-14} = [H^+] \times 0.005$

$$[H^+] = \frac{1.0 \times 10^{-14}}{0.005} = 2.0 \text{No}^{-12} \text{M}$$

As $K_w = [H^+][OH^-]$

$1.0 \times 10^{-14} = [H^+][OH^-]$

Taking log of both sides

$\log (1.0 \times 10^{-14}) = \log [H^+] + \log [OH^-]$

$-14 = -\log [H^+] - \log [OH^-]$

$14 = (-\log [H^+]) + (-\log [OH^-])$

$14 = pH + pOH$

This equation provides another way to express the relationship between the $[H^+]$ and $[OH^-]$ concentrations.

Example 8.2:

Calculate the pH of 0.001 M aqueous hydrochloric acid solution.

Solution:

Hydrochloric acid ionises in water completely therefore,

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

0.001 M 0.001 M

$[H_3O^+]$ is in fact the same as $[H^+]$

Therefore $[H^+] = 0.001$ M

$pH = -\log (0.001)$

$= -\log 10^{-3}$

$= 3.00$

Therefore, the pH of 0.001 M aqueous hydrochloric acid is 3.00

Example 8.3:

Calculate the pH of 0.062 M NaOH solution.

Solution:

$$NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

0.062M 0.062M 0.062M

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

$= -\log [0.062]$

$= 1.21$

Now $pH + pOH = 14$

$pH = 14 - pOH$

$= 14 - 1.21$

$= 12.79$

Self Check Exercise 8.1

What is the pH of a solution containing 1.95g pure H_2SO_4 per dm^3 of solution? (Ans: pH = 1.4)

TITRATION:

"It is defined as a method to find the volume of the standard solution required to react completely with known volume of another solution under analysis".

Acid-base titrations are conducted using burettes and volumetric pipettes. Generally acid solution is placed with the burette. A fixed volume of base is placed into a conical flask along with a few drops of the appropriate acid-base indicator. Acid from the burette is added to the base until the indicator changes colour. This change of colour indicates the end point of titration. In the neutralization of a strong acid with a strong base, phenolphthalein is used as an indicator. It imparts pink colour to the base solution. At the end point, solution just becomes colourless. The molarity of the acid solution under test is determined with the help of the following equation.

$$M_1 V_1 = M_2 V_2$$

Where M_1 = Molarity of the base

V_1 = Volume of the base taken in flask

M_2 = Molarity of the acid

V_2 = Volume of acid used from the burette

Knowing the three parameters, the fourth can be calculated.

Activity 1

Activity for Students

Perform acid-base titrations to calculate molarity and strength of given sample solutions.

Suppose we take strong acid (HCl) and strong base (NaOH) solutions. Molarity of NaOH is known and that of HCl is to be determined. For titration purposes phenolphthalein is used as an indicator. HCl is taken in the burette while 10 ml of NaOH in the conical flask. Add few drops of phenolphthalein in NaOH solution from the Conical flask in it drop by drop till the pinle color just disappears. The disappearance of pink colour gives the end point.

According to molarity concept

$M_1 V_1 = M_2 V_2$

HCl = NaOH

$$M_1 \times 10 = \frac{1}{10} \times 10 \quad \text{or} \quad M_1 = \frac{1}{10} = 0.1M$$

Thus molality of the given solution is 0.1M.

Strength dm^3 = molality x molar mass

$$= \frac{1}{10} \times 36.5 = 3.65\,g\,dm^{-3}$$

Self Check Exercise 8.2

Example 5:

In a titration it is found that 25 cm^3 of 0.12M NaOH is neutralized with 30cm^3 of HCl of unknown concentration. Calculate concentration and strength of HCl solution.

Solution:

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O$$

The Student should be guided to repeat this information in

$n_1 = 1 \quad n_2 = 1$

HCl NaOH

$$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2}$$

$$\frac{M_1 \times 30}{1} = \frac{0.12 \times 25}{1}$$

$$M_1 = 0.1M$$

Thus Molarity of HCl solution is 0.1M.

Strength of solution = Molarity × Molar mass

Strength of HCl solution = 0.1×36.5

$= 3.65 g\ dm^3$

8.4.2 Strong and Weak Acids

The extent of ionization and the acid dissociation constant K_a can be used to distinguish between strong and weak acids.

According to the Bronsted-Lowery definition an acid is a proton donor species. Different acids have a tendency to donate proton but this tendency is different. An acid may donate proton to very high degree, a high degree or to small degree or very small degree. Depending on this ability of donation of proton, an acid is regarded as strong acid or a weak acid. Strong acids donate proton to a greater degree than the weak acids. The strength of an acid is generally expressed in terms of the acid ionization constant, K_a. Consider the case of ionization of a general acid HX in water. In this aqueous solution, the established equilibrium may be represented as follows:

$$HX_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + X^-_{(aq)}$$

The equilibrium constant K for this ionisation process may be written as follows:

$$K = \frac{[H_3O^+][X^-]}{[HX][H_2O]}$$

or $$K[H_2O] = \frac{[H_3O^+][X^-]}{[HX]}$$

Since water is a solvent, it is present in excess and therefore its concentration may be regarded as constant. Thus, $K[H_2O]$ is another constant and is designated as K_a thus,

$$K[H_2O] = K_a = \frac{[H_3O^+][X^-]}{[HX]}$$

K_a is termed as the acid dissociation constant. It is a measure of the extent to which an acid is ionized or dissociated at the equilibrium state. It must be kept in mind that the acid dissociation constant, K_a, is dependent on temperature. Therefore, the value of K_a should be mentioned along with the temperature at which K_a was determined. Dissociation constant, K_a, of

acetic acid in water at 25°C is 1.8×10^{-5}. The comparison of K_a Values of different acids provides a method to compare their strengths.

"The greater the value of K_a, the stronger is the acid".

The value of K_a are usually inconvenient numbers, therefore, for convenience these values are converted to pK_a values. The relationship between K_a and pK_a is as follows:

$$pK_a = - \log K_a$$

Since pK_a refers to the negative logarithm of K_a, smaller the value pK_a stronger shall be the acid because smaller pK_a value corresponds to a greater K_a value. In table 8.2 are listed the ionisation constants and pK_a values of some common acids in water at 25°C. Which acid is strongest acid? Which acid is weakest acid?

Table 8.2: Ionisation constants and pK_a of Acids

Name of Acid	Formula	Ka	pKa
Perchloric acid	$HClO_4$	1.0×10^{10}	-10.0
Hydroiodic acid	HI	1.0×10^{10}	-10.0
Hydrobromic acid	HBr	1.0×10^{9}	-9.0
Hydrochloric acid	HCl	1.0×10^{6}	-6.0
Sulphuric acid	H_2SO_4	1.0×10^{3}	-3.0
Hydrofluoric acid	HF	1.0×10^{-4}	+3.1
Formic acid	HCOOH	1.0×10^{-4}	+3.7
Benzoic acid	C_6H_5COOH	6.3×10^{-5}	+4.2
Acetic acid	CH_3COOH	1.8×10^{-5}	+4.7
Phenol	C_6H_5OH	1.3×10^{-10}	+8.9
Water	H_2O	1.8×10^{-16}	+15.7

Which acid is stronger HCl or HF?

Example 6:

Calculate concentration of H^+ ions of a solution that contains 1.0M HF ($K_a = 7.2 \times 10^{-4}$)

Solution:

	$HF_{(aq)}$	$\rightleftharpoons$	$H^+_{(aq)}$	+	$F^-_{(aq)}$
Initial conc.	1.0M		0		0
Eq. conc. (moles dm^{-3})	1.0 –		X		X

The Student should be guided to distinguish

$$K_a = \frac{[H^+][F^-]}{[HF]}$$

$$7.2\times10^{-4} = \frac{x \cdot x}{1.0-x}$$

Since x is very small as compared to 1.0, the term in the denominator can be approximated as follows:

$$1.0 - x = 1.0$$

$$7.2\times10^{-4} = \frac{x^2}{1}$$

$$x = 0.268M$$

$$[H^+] = 0.268M$$

Example 8.7:

The pH of a 0.1M solution of an acid is 2.85. Calculate the ionization constant, K_a of the acid.

Solution:

Using the pH value, the equilibrium concentration of the acid and the conjugate base can be calculated. These concentrations may then be used to calculate K_a.

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

$$2.85 = -\log [H^+]$$

Taking antilog "of" both side $[H^+]$ turns out to be 1.4×10^{-3}M.

Now equilibrium is considered:

$$Acid_{(aq)} \rightleftharpoons H^+_{(aq)} + \text{conjugate base}_{(aq)}$$

Before equilibrium	0.100M	0.000	0.000
Change	0.0014M	0.0014M	0.0014M
At equilibrium	0.0986M	0.0014M	0.0014M

$$K_a = \frac{[H^+]\,[\text{conjugate base}]}{[\text{acid}]}$$

$$= \frac{(0.0014)(0.0014)}{0.0986}$$

$$= 1.99 \times 10^{-5}$$

Thus the ionisation constant, K_a is 1.99×10^{-5}

8.4.3 Strong and Weak Bases

Use the extent of ionization and the base dissociation constant K_b to distinguish between strong and weak bases.

The strength of a base is the ability to accept a proton from a solvent. Hydroxides of alkali metals such as sodium hydroxide and potassium hydroxide are strong bases and ionize completely in aqueous solution.

$$NaOH_{(aq)} \longrightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$$

$$KOH_{(aq)} \longrightarrow K^{+}_{(aq)} + OH^{-}_{(aq)}$$

The OH^- ion thus formed is a Bronsted base because it can accept proton H^+.

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

The ability of a base to accept a proton from an acid, usually water, is termed as strength of the base. For a base B, an equilibrium reaction with water can be represented by the following equation:

$$B + H_2O_{(aq)} \rightleftharpoons BH^{+}_{(aq)} + OH^{-}_{(aq)}$$

The equilibrium constant K_b is referred to base ionization constant and can be derived in the same way as K_a for acids. Thus:

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

K_b value will be large if degree of ionization of the base B is high i.e. if the base B is strong. The Value of K_b will be small for a weak base B. Again, for convenience, a parameter pK_b has been devised to express K_b value in convenient numbers. Thus, pK_b is defined as the negative logarithm of K_b.

$$pK_b = -\log K_b$$

Table 8.3: K_b and pK_b Values of Some Common Bases

Name of Base	Formula	Kb	pKb
Diethlyamine	$(C_2H_5)_2\,NH$	9.6×10^{-4}	3.02
Ethylamine	$C_2H_5\,NH_2$	5.6×10^{-4}	3.25
Methylamine	$CH_3\,NH_2$	4.5×10^{-4}	3.34
Ammonia	NH_3	1.7×10^{-5}	4.76
Pyridine	$C_5H_5\,N$	5.6×10^{-9}	8.25
Aniline	$C_6H_5\,NH_2$	4.3×10^{-10}	9.37

According to these values ammonia is a stronger base than pyridine and aniline but weaker than methylamine and ethylamine. Also, diethyl amine is a strongest base among all those listed in the table.

8.4.4 Relationship of K_a and K_b:

For a conjugate acid base pair an important relationship exists between K_a, the acid dissociation constant, and K_b, the base dissociation constant. This is given by,

$$K_a \times K_b = K_w$$

The student should be guided to distinguish between weak and strong bases

This relationship can be proved by considering the ionization of an acid HA and writing equation for its K_a

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

$$K_b = \frac{[H^+]\,[A]}{[HA]}$$

The equilibrium for the conjugate base is

$$A^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HA_{(aq)} + OH^-_{(aq)}$$

$$K_b = \frac{[HA]\,[OH^-]}{[A^-]}$$

Multiplying the expression for K_a and K_b.

$$K_a K_b = \frac{[H^+]\,[A]}{[HA]} \times \frac{[HA]\,[OH^-]}{[A^-]}$$

Thus $K_a K_b = [H^+]\,[OH^-] = K_w$

This result leads to an important conclusion that if the value of K_a is known K_b can be calculated since $K_a, K_b, = K_w$

$$\log (K_a K_b) = \log K_w$$

$$\log K_a + \log K_b = \log K_w$$

On changing the signs from plus to minus in the above equation.

$$(-\log K_a) + (-\log K_b) = (-\log K_w)$$

Note that in terms of pK_a & pK_b the equation becomes

$$pK_a + pK_b = pK_w$$

As $pK_w = 14$ at 25°C

$$pK_a + pK_b = 14 \text{ at } 25°C$$

Example 8.8 (a)

The pK_a of acetic acid at 25°C is + 4.76. Calculate the pK_b of the conjugate base of acetic acid.

Solution:

$pK_a + pK_b = 14.00$

pK_a of acetic acid is given as + 4.76

$4.76 + pK_b = 14$ or $pK_b = 9.24$

Example 8.8 (b)

The pK_b of pyridine at 25°C is 8.25. Calculate the pK_a of the

Solution:

$pK_a + pK_b = 14.00$

pK_b of pyridine $= + 8.25$

therefore $pK_a + 8.25 = + 14.00$

and $pK_a = + 14.00 - 8.25$

$= + 5.75$

Therefore, the pK_a of the conjugate acid of pyridine turns out to be +5.75 one of the important conclusion of the equation $K_aK_b = K_w$ is

$$K_a = \frac{K_w}{K_b}$$

and $$K_b = \frac{K_w}{K_a}$$

Since K_w is constant at a given temperature, it may be deduced that K_a is inversely proportional to K_b. Thus, stronger the acid, weaker is its conjugate base. It can also be said that stronger a base, weaker is its conjugate acid. For example, ammonia is a weaker base ($pK_c = + 4.76$) than diethylamine ($pK_b = 3.02$, therefore the conjugate acid of ammonia is stronger acid than the conjugate acid of diethylamine. Similarly, the conjugate base of a weak acid, water $(pK_a = + 6.00)$ is a stronger base than the conjugate base of stronger acid hydrochloric acid $(pK_a = -7.00)$. Thus, hydroxide ion (OH^-) is a stronger base than the chloride ion (Cl).

Self Check Exercise 8.3

Example 9:

What is the percentage ionisation of acetic acid in a solution in which 0.1 M of it has been dissolved per dm^3 of the solution

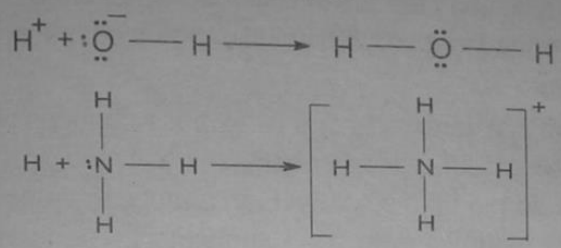
$K_a = 1.8 \times 10^{-5}$

Hint: Percentage ionisation $= \frac{\text{concentration of ionised acid}}{\text{original concentration}} \times 100$

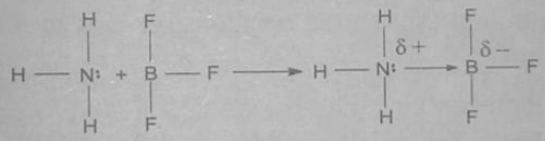
(Ans: 1.3%)

8.5 LEWIS DEFINITIONS OF ACID AND BASE

So far, acid-base concept has been discussed in terms of the Bronsted – Lowery Theory of acids and bases. To behave as a Bronsted base, a substance must be able to accept proton, thus both the hydroxide ion and ammonia are bases:



It should be noted that the hydroxide ion and ammonia, both, have at least one unshared pair of electrons which is used to accept the proton, G.N Lewis, in 1932, put forward his acid-base theory on the basis of this electron pair. According to Lewis definition, a base is a substance that donates a pair of electrons, and an acid is a substance which can accept a pair of electrons. For example, the hydroxide ion (OH^-) is a Lewis base because it donates a pair of electron and the proton (H^+) is a Lewis acid-since it accepts a pair of electrons. The significance of the Lewis acid-base concept is that it is much more general than other concepts. It may include such acid-base reactions which are not covered by the Bronsted–Lowery theory. One such example is the reaction between ammonia and boron trifluoride (BF_3) illustrated by the following equation:



The vacant, unhybridized 2p orbital of boron atom in boron trifluoride accepts the electron pair from ammonia. Thus in this reaction ammonia is a Lewis-base and boron trifluoride is a Lewis acid although no proton transfer is observed here.

In Table 8.4 some common Lewis acids and bases are listed. A careful looks at this table reveals that Lewis acids have the ability to accept electron pair where as the Lewis bases are capable of donating electron pair.

Table 8.4: Some common Lewis Acids and Bases

Lewis Acids		Lewis Bases	
Name	Formula	Name	Formula
Proton	H^+	Hydroxide ion	OH^-
Boron trifluoride	BF_3	Ammonia	NH_3
Aluminium chloride	$Al\,Cl_3$	Carbon monoxide	CO
Silver cation	Ag^+	Water	H_2O
Carbon dioxide	O = C = O	Cyanide ion	$^-C \equiv N$

8.6 BUFFER SOLUTIONS AND THEIR APPLICATIONS (Major Concept)

Multiple industrial processers' rely on buffers, including pharmaceutical ford and to textiles.

A buffer solution is a solution the pH of which does not change significantly when a small amount of acid or base is added to it. Such a solution has a constant pH which does not change on keeping it for a long time.

An aqueous solution of a weak acid or a weak base contains two substances which react with each other to a limited extent. These substances are acid and water or base and water. When a third substance is added to this solution, an effect can be produced on the original pair of reactants in solution (acid and water or base and water). The third substance may suppress the reaction of the original two reactants. For example, if sodium acetate (CH_3COONa) is added to an aqueous solution of acetic acid ($CH_3.COOH$), it does not react directly with acetic acid or with water. However, the increase of the acetate ions suppresses the ionisation of acetic acid. The equation below illustrates the equilibrium of the aqueous acetic acid solution:

$$H_3CCOOH_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COO^{\bullet}_{(aq)} + H_3O^{+}_{(aq)}$$

On addition of sodium acetate (CH_3COONa) the solution receives sodium ions (Na^+) and acetate ions CH_3COO^- as a result of ionisation of sodium acetate. The presence of sodium ion does not affect the equilibrium since it is not involved in this reaction. However, the increase in the concentration of acetate ions (appearing on the right side of the equilibrium equation) will shift the equilibrium to the left.

Now consider a solution containing both a weak acid and its conjugate base (as a soluble salt). Such a solution is called a buffer solution. One buffer solution can be made by dissolving sodium acetate and acetic acid in water. It may be recalled that acetic acid is a weak acid and acetate ion is a conjugate base of acetic acid. Thus in this solution both acetic acid and its conjugate base are present. The acidity of this solution is less than a solution of acetic acid because the acetate ions hinder the ionisation of acetic acid. In addition, the buffer solution has a special character of resisting change in its pH even when a small amount of strong acid or strong base is added to it. Such a solution has the ability to neutralize either added acid or a base. If base is added to the buffer solution, the hydroxide ions (OH^-) will be neutralized by the acid (acetic acid) in the buffer.

$$H_3CCOOH_{(aq)} + OH^-_{(aq)} \longrightarrow CH_3COO^-_{(aq)} + H_2O_{(l)}$$

On the other hand if an acid (say HCl) is added, the protons (H^+ ion) produced will be consumed by the acetate ions present in the buffer:

$$H_3CCOO^-_{(aq)} + H^+_{(aq)} \longrightarrow CH_3COOH_{(aq)}$$

A buffered solution may contain a weak acid and its salt from strong base (e.g. HF and NaF) or weak base and its salt from strong acid (e.g. NH_3 and NH_4Cl). By choosing the appropriate components, a solution can be buffered at

The following example explains the calculations associated with buffered solutions.

The concentration of conjugate base in the reaction mixture is predominately supplied by the salt which is a strong electrolyte. Therefore, assuming the concentration of conjugate base equal to that of salt and original concentration of acid as equilibrium concentrations, pH of a buffer can be calculated.

Science Titbits

Fermentation reactions such as in breads the pH of dough drop with the production of CO_2 and some organic acids. In this case both natural buffers milk and flour, as well as sodium hydrogen carbonate can be used to limited pH variations during fermentation.

Example 8.9:

(a) Calculate the pH of an acetic acid-sodium acetate buffer solution containing 1.0 moles of each component.

(b) What will be the pH of this solution after addition of 0.01 mole of hydrochloric acid gas to 1dm^3 volume? Assume that the volume of solution remains unchanged on addition of hydrochloric acid. (Ka for acetic acid is 1.8×10^{-5}.

(c) The pH of the buffer solutions can be calculated by assuming the equilibrium concentration of both the acid and its conjugate base as starting concentration.

Thus $[CH_3CO_2H] = 1.0M$ $[CH_3COO] = 1.0M$

For acetic acid dissociation

$$K_a = \frac{[CH_3COO][H^+]}{[CH_3CO_2H]} = 1.8 \times 10^{-5}$$

$$1.8 \times 10^{-5} \frac{[1.0][H^+]}{[1.0]}$$

Thus, the pH of buffer solution is 4.745.

(d) After HCl addition:

$$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^{-1}_{(aq)}$$

0.01 mole 0.01 mole 0.01 mole

Initially, there were 1.0 mole of CH_3COOH and 1.0 mole of CH_3COO^- present per dm^3 of the solution. After addition of hydrochloric acid 0.01 mole of CH_3COO^- ions are combined with the H^+ ions formed from dissociation of 0.01 mole of added hydrochloric acid. This can be written as:

$$CH_3COO^-_{(aq)} + H^+_{(aq)} \longrightarrow CH_3CO_2H_{(aq)}$$

0.01 mole 0.01 mole 0.01 mole

Thus, the numbers of moles of acetic acid and acetate ions, after addition of hydrochloric acid are:

CH_3CO_2H : (1.0+0.01) mole = 1.01mole

CH_3COO^- : (1.0-0.01) mole = 0.99mole

The equilibrium equation for this new situation can be written

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3CO_2H]}$$

or $$[H^+] = \frac{Ka[CH_3CO_2H]}{[CH_3.COO^-]}$$

$$[H^+] = \frac{1.8 \times 10^{-5}(1.01)}{0.99}$$

$$pH = -\log (2.210^{-5})$$

$$= 4.736$$

Notice there is a slight change in pH from 4.745 to 4.736 that is only a difference of 0.009. Thus a buffer does a very good job in limiting the change in pH to a very small amount. Some important facts about buffer solutions are as follows:

1. Buffer solution can be prepared by combining, in aqueous solution, a weak acid and its salt with strong base or a weak base and its salt with strong acid.
2. A buffer solution resists change in its pH even if a small amount of strong acid or a base is added to it.
3. A buffer solution maintains the stability of its pH by shifting their equilibrium to consume added H^+ ion or to replace H_3O^+ ions which have reacted with the added OH^- ions.
4. An aqueous solution of a strong acid and its conjugate base (as its soluble salt) can not act as a buffer solution. Since the strong acid is completely ionized already, addition of small amount of base (OH^- ions) will consume the H_3O^+ ions which can not be replaced and the pH of the solution will change. If a small amount of strong acid is added to this solution, the equilibrium will not shift to the left because the acid already present is a strong acid. Also a solution of strong base and its conjugate acid do not form a buffer solution due to similar reasons.
5. Some common examples of buffers are acetic acid/sodium acetate buffer, phosphoric acid/potassium dihydrogen phosphate buffer and formic acid/sodium format buffer.

Applications of buffer solutions:

(1) Buffer solutions play an important role in several industrial processes. For example they are used in the manufacture of photographic materials, leather and dyes

(2) They are also used in the process of electroplating and analytical procedures.

(3) The buffer solutions are also used for calibration of pH meters.

(4) In nature, many biological systems depends upon buffer action to preserve a constant pH e.g. the pH of human blood is maintained between 7.35 to 7.45. The pH of tears is also maintained at 7.40 whereas the stomach juices are known to have pH preserved at 1.65 – 1.75. Milk and egg white (protein) show their pH as 6.7 – 6.8 and 8.0 – 8.1 respectively.

(5) Buffers are used to maintain the pH of culture media for the growth of bacteria in bacteriological applications.

Composition of a buffer solution:

A buffer solution can be made in two ways:

(1) By mixing a weak acid and a salt of it with a strong base. Such solutions give acidic buffers with pH less than 7. e.g. $CH_3COOH + CH_3COONa$.

(2) By mixing a weak base and a salt of it with a strong acid. Such solutions will give basic buffers with pH more than 7. e.g. $NH_4OH + NH_4Cl$.

Buffer Action:

(i) Let us take a buffer solution of CH_3COOH and CH_3COONa. Common ion effect helps us to understand how will buffer work. CH_3COOH being a weak electrolyte undergoes very little dissociation. When CH_3COONa, a strong electrolyte is added to CH_3COOH solution, the dissociation of CH_3COOH is suppressed due to common ion effect of CH_3COO^-.

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

$$CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$$

Suppose we add a few drops of HCl to it. Its H^+ ions are used up by CH_3COO^-.

Thus the addition of HCl will not change the pH of the buffer solution.

(ii) In the same buffer solution, if a strong base is added it is neutralised by the acid.

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$

$$NaOH \rightleftharpoons Na^+ + OH^-$$

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

Thus the addition of NaOH will not change value of pH.

Example 8.10:

What is the pH of buffer if concentration of CH_3COOH is 0.1 M and CH_3COONa is 1.0 M pK_a for CH_3COOH is 4.76?

Solution:

Concentration of CH_3COOH = 0.1 M

Concentration of CH_3COOHa = 1.0 M

The formula to determine pH value is

$$pH = pK_a + \log \frac{[salt](i.e.\,CH_3COONa)}{[acid](i.e.\,CH_3COOH)}$$

$$pH = 4.76 + \log \left(\text{since } \frac{1.0}{0.2} = \frac{10}{1} \right)$$

$$pH = 4.76 + \log 10 \text{ (as } \log 10 = 1)$$

$$pH = 4.76 + 1.00$$

$$pH = 5.76$$

Self Check Exercise 8.4

Calculate the pH of a buffer solution in which 0.11 Molar CH_3COONa and 0.09 Molar CH_3COOH solutions are present.

Ka for CH_3COOH is 1.8×10^{-5} **(Ans:** 4.83)

8.7 SALT HYDROLYSIS

Consider the following observations:

(i) Aqueous solution of NH_4Cl turns blue litmus red.

(ii) Aqueous solution of K_2CO_3 turns red litmus blue.

(iii) Aqueous solution of NaCl has no action on litmus solutions.

These observations can be explained on the basis of Bronsted-Lowry acid-base theory. When a salt (MX) is dissolved in water, it splits up into its M^+ and X^- ions. These ions may react with water and give following reactions:

$$M^+ + H\text{-}OH_{(\ell)} \rightleftharpoons MOH_{(aq)} + H^+_{(aq)}$$

$$\overline{X}_{(aq)} + H\text{-}OH_{(\ell)} \rightleftharpoons HX_{(aq)} + OH^-_{(aq)}$$

Since H^+ and OH^- ions are produced in these reactions, the solution of the salt may be acidic or basic. In salts anions are derived from acids and cations from bases. The anions of weak acids are strong conjugate bases. Such anions react with water producing basic solutions

For example: CH_3COO^{-1}, CN^{-1}, HCO^{-1}_3, CO^{-2}_3 etc.

$$CH_3COO^{-1}_{(aq)} + H\text{-}OH_{(\ell)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$$

$$CN^-_{(aq)} + H\text{-}OH_{(\ell)} \rightleftharpoons HCN_{(aq)} + OH^-_{(aq)}$$

$$CO^{-2}_{3\,(aq)} + 2H\text{-}OH_{(\ell)} \rightleftharpoons H_2CO_{3\,(aq)} + 2OH^-_{(aq)}$$

The student should be guided to

Anions like Cl^-, NO_3^-, SO_4^{-2}, are so weak conjugate bases that, they do not react with water. Cations of weak bases are strong conjugate acids. Such cations react with water producing acidic solutions.

For example: Cu^{+2}, Al^{+3}, NH_4^+, etc.

$$Cu^{+2}_{(aq)} + 2H - OH_{(\ell)} \rightleftharpoons Cu(OH)_{2(aq)} + 2H^+_{(aq)}$$

$$NH^+_{4(aq)} + H - OH_{(\ell)} \rightleftharpoons NH_{3(aq)} + H_3O^+_{(aq)}$$

Do You Know

. Hydrolysis is an important process in plants and animals. In living systems, most, biological reactions including ATP hydrolysis, take place during the catalysis of enzymes. The catalytic action of enzymes allows the hydrolysis of fets dils, proteins and carbohydrates.

Cations like Na^+, K^+, Ca^{+2}, Na^+, K^+, Ca^{+2}, Mg^{+2}, etc. are so weak conjugate acids that they do not react with water.

These reactions are called hydrolysis reactions. "The reaction of cations and anions of salts with water is called hydrolysis."

Hydrolysis is different from hydration. In hydrolysis H – OH bond is broken whereas in hydration water molecule adds up to a substance without bond breakage.

There are four types of salts on the basis of their reactivity with water:

1. Salts of strong acids and strong bases do not hydrolyse (pH = 7).

Examples: NaCl, Na_2SO_4, KNO_3 etc.

2. Salts of weak acids and strong bases hydrolyse producing basis solutions (pH>7).

Examples: CH_3COONa, NaCN, Na_2S etc.

3. Salts of strong acids and weak bases hydrolyse producing acidic solutions (pH < 7).

Examples: $CuSO_4$, NH_4Cl, NH_4NO_3 etc.

4. Salts of weak acids and weak bases hydrolyse, but the resulting solution is either neutral, acidic or basic. This depends upon the relative values of K_a and K_b of cations and anions of the salt.

The important aspects of the salt hydrolysis, discussed above, are summarised in Table 8.5.

8.5: Summarised discussed as above

Salt Type		Common Example	Ions which impart Hydrolysis	Solution pH (Nature)
Acid	Base			
Strong	Strong	NaCl, K Br	None	= 7.0 (Neutral)
Strong	Weak	NH_4NO_3, NH_4Cl	Cations	< 7.0 (Acidic)
Weak	Strong	NaCN, K_2CO_3	Anions	> 7.0 (Basic)
Weak	Weak	NH_4CN, NH_4NO_2	Anions & Cations	May be equal, smaller or greater than 7.0

The student should be guided to discuss and explain the concept.

8.73 The Levelling Effect

One way to compare the strength of acids is to select a reference base (usually the solvent) and determine the extent of proton donation from the acids to the reference base. However, the reaction between any stronger acid than hydroxonium ion (H_3O^+) and water goes completely to the right. Strong acids like perchloric acid, hydrochloric acid, nitric acid and Sulphuric acid in water appear to be of equal strength. Thus, water as a base, is unable to differentiate among the relative acid strength of acids stronger than hydroxonium ion. Such inability of any solvent to differentiate among the relative strength of all acids is termed as the levelling effect. Because the solvent is said to level the strength of all the acids, making them appear identical.

The levelling effect of water can be compensated for, if a more weakly basic solvent like acetic acid is employed in place of water. Since acetic acid is a much weaker base than water, it is not easily protonated. Thus, appreciable differences in proton donation of acids are observed in acetic acid (solvent).

The relative acid strength is $HI > HBr > HCl > H_2SO_4 > HNO_3$

It may be pointed out that all these acids are of identical strength in water (solvent) due to the levelling effect.

e.g. $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$

$HI + H_2O \longrightarrow H_3O^+ + I^-$

$HBr + H_2O \longrightarrow H_3O^+ + Br^-$

$H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO_4^-$

$HNO_3 + H_2O \longrightarrow H_3O^+ + NO_3^-$

Hydroxonium ion (H_3O^+), which is acidic in nature, is formed in each case. The equal behaviour of such ions formed levels the strength of the acid.

References for additional information

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- John W. Hill & Synthia S. Hill, R. D., Chemistry for changing times.
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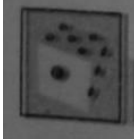

Exercise

1 Encircle the correct answer in each case:

i. Water cannot act as:

(a) Lewis acid
(b) Lewis base
(c) Bronsted acid
(d) Bronsted base

ii. pH of 0.01M HCl solution is:
(a) 10^{-2} (b) 10^{+2} (c) 2.0 (d) 1.0

iii. An aqueous solution of ammonium chloride is:
(a) Basic (b) Acidic (c) Neutral (d) Amphotric

iv. An aqueous solution of which compound is basic:
(a) Ammonium nitrate (b) Calcium chloride
(c) Ammonium acetate (d) Potassium carbonate

v. Which statement about acids is not correct? An acid:
(a) Contains hydrogen ions in solution.
(b) Contains oxygen.
(c) Has a pH of less than 7.
(d) Gives off carbon dioxide from a carbonate.

vi. If a liquid has a pH of 7.
(a) It must be colourless.
(b) It has boiling point of 100°C.
(c) It must be a solution.
(d) It must be neutral.

$CO_2 + H_2O \longrightarrow H_2CO_3$

vii. When air is bubbled through pure water, the pH is lowered from 7.0 to 5.6, which gas in the air is responsible for this change.
(a) Argon (b) Carbon dioxide (c) Nitrogen (d) Oxygen.

viii. Which of the following oxides is classified incorrectly:
(a) Zinc oxide (ZnO) amphoteric
(b) Carbon dioxide (CO_2) acidic
(c) Carbon monoxide neutral
(d) Aluminium oxide (Al_2O_3) basic

ix. If $25cm^3$ of 1 $mol.dm^{-3}$ nitric acid is added to $50cm^3$ of 0.5 potassium hydroxide solution, what would be the pH of the resulting solution
(a) 5 (b) 7 (c) 9 (d) 14

x. If dry citric acid crystals are placed on dry litmus paper they will
(a) turn yellow (b) turn green
(c) turn red (d) remains unchanged

xi. A base is a substance which will neutralise an acid which of these substances is not a base:
(a) Aqueous ammonia (b) Copper oxide
(c) Potasium chloride (d) Sodium carbonate

xii. A strong acid:
(a) Is always partially ionised when in solution.
(b) Is always fully ionised when in solution.
(c) Always decomposes carbonates.
(d) Always contains oxygen.

xiii. Which one of the following oxides dissolves in water to form acidic solution:
(a) MgO (b) Na_2O (c) SO_2 (d) SiO_2

xiv. When crystals of copper sulphate are heated, the colour changes from blue to white. This is caused by:
(a) Loss of water only
(b) Loss of water and SO_2.
(c) Reaction with CO_2 in the air.
(d) Loss of water, sulphur dioxide and oxygen.

xv. The oxide of a metal was found to react with HCl and aqueous NaOH solution. Which of the following is the best description of the oxide?
(a) Acidic (b) Amphoteric (c) Basic (d) Neutral

SHORT QUESTIONS

(i) What are acidic, basic and amphoteric substances? Give one example of each substance.

(ii) Elaborate with equations two compounds which are amphoteric in nature.

(iii) What is Bronsted - Lowry acid-base theory? Give examples.

(iv) What are conjugate acid-base pairs? Explain with examples.

(v) Define Lewis acid and Lewis bases. Give one example in each case.

(vi) Classify each of the following as Bronsted acid or Bronsted base.
(i) HCO_3^{-1} (ii) HBr (iii) CH_3COO^{-1}

(Vii) Explain gastric acidity and use of anti-acid drug.

(Viii) Write briefly about the ionisation of water.

(ix) Define pH what are the values of pH for acidic, basic and neutral solutions.

(x) Define K_a and pKa and their applications.

(xi) Explain curdling of milk with lemon juice.

(xii) What are Kb and pK_a and their applications?

(xiii) What is an iodised salt? What is its use in practical life?

(xiv) What is the relationship between K_a and K_b?

(xv) Calculate concentrations of ions of slightly soluble salts.

(Xvi) Give two examples of a buffer solution.

Elaborate the ionization equation of water. How does it lead to the ion-product constant of water? Describe an important application of the ion-product of water.

What are buffer solutions? Elaborate with suitable examples, their significance in acid-base reactions. Write three common applications of buffer solutions.

Write detailed notes on each of the followings:
(a) Conjugate acid base pairs (b) pK_a (c) pK_b (d) Relative strength of acids.

6 What is hydrolysis? Discuss in detail, the behaviour of each of the following salts in their aqueous solutions.

(a) K_2CO_3 (b) NH_4Cl (c) $NaNO_3$

7 (a) Calculate the pH of formic acid-sodium format buffer solution containing 1.0 mole of each component.

(**Ans**: 3.7447)

(b) What will be the pH of the solution after addition of 0.10 mole of hydrochloric acid gas to $1.01dm^3$ volume of the ,buffer solution in part (a) Assume that the volume of solution remains unchanged on addition of hydrochloric acid (K_a for formic acid is 1.8 x 104)

(**Ans**: 3.6997)

8 (a) Calculate the H^+ ion concentration of an aqueous solution having pH 10.6

(**Ans**: 2.5×10^{-11} moles/dm^3)

(b) An aqueous solution contain 1.0×10^{-9} moles/dm^3 of hydronium ions. Calculate the pOH of this solution

(**Ans**: 5.0)

9 (a) What is acid dissociation constant? How is it related to pK_a? Write equation to elaborate.

(b) Define and briefly describe the levelling effect of water in acid-base reactions.

10 (a) What is hydrolysis? Write the equations of hydrolysis equilibrium for each of the followings:

(i) Li^+ (ii) NH^{+4} (iii) CN^-

(b) Write a unique property of a buffer solution. What types of acids are required to prepare buffer solutions? Name two such acids.

(vi) $\rightarrow H_2CO_3 + OH$